Elementary Chemistry

THE ATOMIC THEORY—CHEMICAL COMBINATION—COMBINING VOLUMES OF GASES—MOLECULAR WEIGHTS OF GASES—FLAME AND THE COMBUSTION OF HYDRO-CARBONS—LIQUIDS—MATTER—PROPERTIES OF GASES—PROPERTIES OF SOLUTIONS — VOLUME OF GASES —THE ELEMENTS

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ILLUSTRATED

CHICAGO

FREDERICK J. DRAKE & CO.

PUBLISHERS

1907

PREFACE

It has been the object of the compiler of this book to present the most important facts and principles of Chemistry in a plain and intelligent manner, as suited to the requirements of students and others who may wish to acquire a knowledge of Elementary Chemistry. It contains full and explicit information on the following subjects: The Atomic Theory—Chemical Combination—Combining Volumes of Gases—Determination of Molecular Weights of Gases—Flame and Combustion of Hydro-Carbons — Liquids—Matter—Properties of Gases—Properties of Solutions—Volume of Gases—The Elements.

THE AUTHOR.

ELEMENTARY CHEMISTRY.

ATOMIC THEORY.

According to the atomic theory, combination takes place between the atoms of which substances are made up, and it hence follows that the number of atoms which is contained in a given volume of any gaseous body must stand in a simple relation to that contained in the same volume of any other gas measured under equal circumstances of temperature and pressure. The simplest as well as the most probable supposition respecting this question is that equal volumes of all the different gases, both elementary and compound, contain the same number of particles or molecules, and this theory is now generally accepted by physicists, who have arrived at the same conclusion as the chemists have reached by an independent train of reasoning. If we take the simplest case of volume combination, that of one volume (one molecule) of chlorine and one volume (one molecule) of hydrogen uniting to form two volumes (two molecules) of hydrochloric acid gas, it is clear that, since each molecule of hydrochloric acid contains at least one atom of chlorine and one of hydrogen, there are at least twice as many atoms as molecules of these elements present. Hence, the molecule of free chlorine and of free hydrogen must consist of at least two atoms com-

bined together, and we shall represent the combination as taking place between one volume (one molecule of two atoms) of chlorine and one volume (one molecule of two atoms) of hydrogen, forming two volumes (two molecules) of the compound hydrochloric acid gas. Again, two volumes of steam are formed from two volumes of hydrogen and one volume of oxygen, hence if there are the same number of molecules of steam, of hydrogen, and of oxygen in the same volume of each gas, it is clear that in the formation of water from its elements, each molecule of oxygen must be split up into two similar parts. We are thus led to distinguish between the atom and the molecule, the latter term being applied to the smallest particle of an element or compound which can exist in the free state. That is, equal volumes of all gases, measured under the same physical conditions, contain equal numbers of molecules.

An immediate consequence of this theory is of the utmost importance. If we weigh equal volumes of two gases, we are obviously weighing equal numbers of their molecules, and the ratio of the weights of the gases will also be the ratio of the weights of their molecules, so that we are thus enabled to determine the relative weights of the molecules of all gases, by simply finding their relative densities. Hydrogen gas is taken as the standard of comparison because it has a lower density than any other gas, and, since it has been shown that its molecule can be divided into at least two parts, the weight of its molecule is taken as equal to two. The molecular weight of any gas is therefore equal to twice its density compared with hydrogen; nitrogen, for example, is about 14 times as

heavy as hydrogen and hence its molecular weight is about 28, whilst carbon dioxide has a density of about 22 and therefore a molecular weight of about 44.

When the molecular weight of a gas and also its composition, as determined by analysis, are both known, it is possible to calculate what proportion of each of the component elements is present in the molecule. Water for instance contains 88.81 per cent of oxygen and 11.19 of hydrogen, whilst the density of steam is about 8.94, its molecular weight being, therefore, equal to 17.88. If now we calculate how much oxygen and hydrogen are present in 17.88 parts of water, we find that this amount is made up of 15.88 of oxygen and 2 of hydrogen. Carbon dioxide again has a molecular weight of 43.7, and contains

Carbon	27.27
Oxygen	72.73
	100.00

Hence 43.7 parts of this gas contain 11.91 of carbon and 31.76 of oxygen.

A repetition of this process for all the known compounds of some particular element enables us to ascertain the least amount of that element which is ever found in a molecule of a substance, and to this amount the name of atom is given. A comparison of all the substances containing oxygen, for example, teaches us that the least amount of it ever found in a molecule is 15.88 parts, and this is, therefore, taken as the atomic weight of oxygen. This having been ascertained, we are in a position to say that the molecule of water contains one atom of oxygen, whilst that of

carbon dioxide contains two. All the non-metallic elements form compounds which can exist in the state of gas, and hence the atomic weights of all these elements have been found by this method. Many of the metals on the other hand do not form volatile compounds, and the atomic weights of these have, therefore, to be determined by different methods. In any case it must be remembered that the method described above is not generally capable of great accuracy and only yields an approximate number for the atomic weight, the exact value being found by determining the equivalent of the element by an accurate analysis of one of its compounds, and then taking as the exact atomic weight the multiple of this number which approaches most closely to the approximate number obtained from the molecular weights.

The ratio of the atomic weights of hydrogen and oxygen has been carefully determined by several investigators with agreeing results, and hydrogen has been kept as the standard and unit, all other atomic weights being calculated from the ratio $O : H = 15.88 : 1$, found by recent experiments.

It will be seen that the determination of the atomic weight is quite distinct from that of the molecular weight. This is well shown in the case of carbon; a comparison of the gaseous compounds of carbon shows that the atomic weight of this element is about 12, this being the least amount of it which is found in the molecule of one of its compounds, we are, however, quite ignorant of the molecular weight of carbon itself, since its density in the state of gas has never been determined.

The molecules of some elements contain as many as

four atoms, phosphorus and arsenic, others contain only two atoms, this being the case with hydrogen, oxygen, nitrogen, chlorine and others, whilst the molecules of the gases of mercury vapor and of the vapors of some of the other metals consist of single atoms, the molecular and atomic weights being, therefore, identical.

For the first time we may now employ chemical symbols, a kind of shorthand, by which we can conveniently express the various chemical changes. To each element is given a symbol, usually the first letter of the Latin, which is generally also that of the English name. Thus O stands for oxygen, H for hydrogen, S for sulphur, Au for gold (aurum), Ag for silver (argentum). These letters, however, signify more than that a particular substance takes part in the reaction. They serve also to give the quantity by weight in which it is present. Thus O does not stand for any quantity, but for 15.88 parts by weight (the atomic weight) of oxygen, H always stands for one part of weight of hydrogen, and in like manner S, Au, and Ag stand invariably for 31.83, 195.7, and 107.12 parts by weight of the several elements respectively. By placing symbols of any elements side by side, a combination of the elements is signified, thus:

HCl=Hydrochloric acid.	HI=Hydriodic acid.
HgO=Mercuric oxide.	HBr=Hydrobromic acid.

If the molecule contains more than one atom of any element, this is indicated by placing a small number below the symbol of the atom of the element, thus H_2O signifies 17.88 parts by weight of a compound (water) containing two atoms or 2 parts by weight of hydro-

gen and one atom or 15.88 parts by weight of oxygen. In such a case as this, where the molecular weight and the number of atoms in the molecule are known and expressed in the formula, the latter is said to be a molecular formula, and consequently represents such a weight of the substance as will in the state of gas occupy the same volume as two parts by weight of hydrogen. When the molecular weight of the compound to be represented by a formula is not known, it is only possible to express the relative number of the atoms of the constituent elements which are present. A formula of this kind is known as an empirical formula and may be calculated for any substance of which the composition has been determined by analysis.

The gas known as ethylene has the following composition as determined by analysis:

Carbon	85.62
Hydrogen	14.38
	100.00

In order to find the empirical formula of this substance it is only necessary to divide the percentage of each element by the atomic weight of the element, which gives us the ratio of the number of atoms of each of the two elements, and if we express this ratio in the smallest possible whole numbers we have at once the relative numbers of atoms present in the molecule, without, however, having any information as to the absolute number.

	Percentage.	$\frac{\text{Percentage}}{\text{Atomic Weight.}}$	Simplest Ratio.
Carbon	85.62	7.19	1
Hydrogen	14.38	14.38	2

The simplest or empirical formula of ethylene is therefore CH_2. The density of this gas, however, is found to be equal to 13.91 and its molecular weight is therefore 27.82, its molecular formula being consequently C_2H_4.

It is usual to represent chemical changes in the form of equations, the materials taking part in the change being placed on one side and the products formed, which are always equal to them in weight, being placed on the other. If we heat potassium chlorate, a substance which has the empirical formula $KClO_3$, it is decomposed into oxygen and potassium chloride and this decomposition is represented by the equation

$$2KClO_3 = 2KCl + 3O_2$$

in which the sign $+$ connects the two products and signifies **together with.** This equation is an expression of the fact, ascertained by experiment, that $243.36 = (38.86 + 35.18 + 3 \times 15.88) \times 2$ parts of this salt by weight leave behind on heating 148.08 parts $= (35.18 + 38.86) \times 2$ of potassium chloride and liberate 95.28 parts of oxygen. Hence it is clear that the quantity of oxygen which is obtained from any other weight of the salt and vice versa can be found by a simple calculation when the equation representing the chemical change is known.

To take a more complicated case, when we know that the equation representing the change which occurs when we heat potassium ferrocyanide, the empirical formula of which is $K_4C_6N_6Fe$, with strong sulphuric acid, H_2SO_4, and water, is the following:

$$K_4C_6N_6Fe + 6H_2SO_4 + 6H_2O = 6CO + 2K_2SO_4 + 3(NH_4)_2SO_4 + FeSO_4$$

yielding carbon monoxide gas CO, potassium sulphate K_2SO_4, ammonium sulphate $(NH_4)_2SO_4$, and iron sulphate $FeSO_4$, we can easily calculate how many grams of carbon monoxide gas, CO, can be obtained from any given weight of the ferrocyanide, $K_4C_6N_6Fe$, inasmuch as analysis proves that the amounts represented by these formula are made up as follows:

Carbon Monoxide.		
Carbon	C	11.91
Oxygen	O	15.88
		27.79

Ferrocyanide of Potassium.		
Potassium	K_4	155.44
Carbon	C_6	71.46
Nitrogen	N_6	83.58
Iron	Fe	55.50
		365.98

The foregoing equation then shows that 365.98 parts by weight of the ferrocyanide yield 166.74 parts by weight of carbon monoxide, and hence a simple proportion gives the quantity yielded by any other weight. The illustration is, however, not yet complete, commercial potassium ferrocyanide contains, as do many crystalline compounds, a certain quantity of water of crystallization, which is given off when the salt is heated, in consequence of which the crystals fall to a powder. But, as the equation shows, a certain quantity of water takes part in the reaction, and it is, therefore, unnecessary to dry the salt previously if only we know how much water of crystallization it contains. Analysis has shown that the commercial salt has the composition $K_4C_6N_6Fe+3H_2O$, hence if we add 3×17.88, the weight of 3 molecules of water, to 365.98, we obtain the number 419.62 as the weight of

the hydrated salt which must be taken in order to obtain 166.74 parts by weight of carbon monoxide.

As, however, the quantity of a gas is almost always estimated by measuring its volume, and from this volume calculating its weight, it becomes of the greatest importance to know how to calculate the volume of a gas from its weight, or vice versa. This can only be effected with strict accuracy by employing in each case the density of the gas as determined by experiment. An approximate number, which it is often useful to know, can, however, be readily obtained, since we know that molecular proportions of all gases occupy equal volumes under the same physical conditions.

Now 1 litre of hydrogen at 0°C. and 760 mm. pressure (which are generally taken as the standard temperature and pressure) weighs 0.089901 gram, and hence the volume occupied by 2 grams, or 1 gram-molecule of the gas is $2/0.089901 = 22.247$ litres. The gram-molecule of every other gas therefore occupies approximately this same volume, or in other words the molecular volume of all gases is 22.247 litres at 0°C. and 760 mm. pressure.

The volume at the standard temperature and pressure of any weight of gas can then readily be calculated, thus 27.79 grams of carbon monoxide occupy 22.247 litres, and hence 166.74 grams occupy

$$\frac{22.247 \times 166.74}{27.79} \text{ litres at } 0°\text{C. and } 760 \text{ mm.}$$

It is now easy to calculate what volume this weight will occupy at any other temperature or pressure, for we know that all gases expand by $\frac{1}{273}$ of their volume at 0°C. when their temperature is raised 1°C. at

constant pressure, and that their volume is inversely proportional to the pressure to which they are subjected. Hence if the temperature at which the gas was collected were 17°C., and if the barometer then stood at 750 mm., the volume v in litres of the carbon monoxide collected would be

$$v = \frac{22.247 \times 166.74 \times (273 + 17) \times 760}{27.79 \times 273 \times 750}$$

CHEMICAL COMBINATION.

The composition of a chemical compound can be ascertained in two ways: By separating it into its component elements, an operation termed **analysis**, and by bringing the component elements under conditions favorable to combination, an operation termed **synthesis.** In both of these operations the scale is employed, the weight of the compound and of the components in each instance must be ascertained, except indeed in the case of certain gases of known specific gravity, when a measurement of the volume occupied by the gas may be substituted for a determination of its weight.

It is one of the aims of **analytical chemistry** to ascertain with great precision the composition of all chemical substances, and this branch of inquiry is termed **quantitative analysis**, as distinguished from that which has only to investigate the kind of material of which substances are composed, and which is hence termed **qualitative analysis.**

Combination by Weight. The first law discovered by the use of the scale, is that the elements combine with one another in a limited number of definite proportions, this number being almost invariably found by experiment to be a small one. When two elements are brought together under such conditions that they can combine, it is always found that one or more of a small number of compounds is produced, the particular substance or substances formed depending upon

the special circumstances of the experiment. Thus carbon is found to be capable of uniting with oxygen in two different proportions, producing two distinct substances, carbonic acid gas and carbon monoxide, these being the only compounds of carbon with oxygen which are known. Some elements, on the other hand, only form one compound with each other, whilst others again form a large number. Each one of these compounds is found to have a fixed composition, containing the elements of which it is made up in a definite proportion by weight, and this fixity of composition is used as a characteristic of a chemical compound as opposed to a mere mechanical mixture, the constituents of which may be present in any variable proportions. In whatever way the conditions under which the elements are made to combine may be varied, it is always found that they unite in exactly the same ratio, unless, as sometimes happen, the changed conditions are favorable to the production of one of the small number of other compounds which can be formed by the same elements. Thus, for instance, the combination of silver with chlorine has been brought about in no less than four different ways, but in every case it has been found that the resulting compound contains 107.12 parts of silver for 35.18 of chlorine. The combination of chlorine with phosphorus, on the other hand, takes place in two distinct ratios, so that when an excess of phosphorus is present, the resulting compound contains 10.25 parts of this element for 35.18 parts of chlorine, whilst if the latter be kept in excess, this weight of it only combines with 6.15 parts of phosphorus. These are, however, the only two compounds of these elements which

are known. In like manner hydrogen combines with oxygen to yield water, a substance which contains 88.81 parts of oxygen to 11.19 of hydrogen. If these elements are brought together in proportions differing from those in which they are present in water, the excess of one element remains in the free state, thus, if 98.81 parts of oxygen by weight be brought together with 11.19 parts of hydrogen under circumstances in which they can combine, 88.81 parts of the oxygen will combine with all the hydrogen to form 100 parts of water, whilst 10 parts of oxygen remain in the free state.

It will therefore be seen that the chemical combination of two or more elements does not result in the production of a series of compounds varying gradually in composition, according to the conditions of the experiment, but yields one or more compounds, each of which contains its constituents in a perfectly fixed and definite ratio.

As has been said, the case frequently occurs of two elements uniting to form several compounds, for each of which the law of definite proportion holds good. Thus the two elements, carbon and oxygen, unite to form two distinct compounds, carbonic oxide gas and carbonic acid gas, and 100 parts of each of these bodies are found by analysis to contain the following weights of the elements:

	Carbonic Oxide Gas.	Carbonic Acid Gas.
Carbon...........	42.86	27.27
Oxygen..........	57.14	72.73
	100.00	100.00

Knowing these facts what is the relation of one element (say of the oxygen) in both compounds when the other element remains constant? In proportion to the carbon, the one compound contains exactly double the quantity of oxygen which the other contains, thus:

	Carbonic Oxide Gas.	Carbonic Acid Gas.
Carbon	10.0	10.0
Oxygen	13.3	26.6
	23.3	36.6

Analysis shows that two compounds which carbon forms with hydrogen, marsh gas and olefiant gas, have the following percentage composition:

	Marsh Gas.	Olefiant Gas.
Carbon	74.95	85.68
Hydrogen	25.05	14.32
	100.00	100.00

How much hydrogen is combined in each compound with 10 parts by weight of carbon. In olefiant gas there are 1.67 parts by weight of hydrogen to 10 of carbon, whilst marsh gas contains 3.34 parts of hydrogen to the same quantity of carbon, or exactly double as much.

As another example we may take the compounds of nitrogen and oxygen, of which no less than five are known to exist. The percentage composition of these five bodies is found by experiment to be as follows:

Nitrogen	63.71	46.75	36.91	30.51	25.99
Oxygen	36.29	53.25	63.09	69.49	74.01
	100.00	100.00	100.00	100.00	100.00

If we inquire how much oxygen is contained in each of these five compounds, combined with a fixed weight, say 10 parts of nitrogen, we find that this is represented by the numbers 5.7, 11.4, 17.1, 22.8, and 28.5. In other words, the relative quantities of oxygen are in the ratio of the simple numbers 1, 2, 3, 4, and 5.

The above examples illustrate the relations exhibited in the combination of two or more of the elements to form compounds, but a careful examination of the quantitative composition of a whole series of chemical compounds leads to a further conclusion respecting the nature of the laws of chemical combination which is of the highest importance. Let us examine the composition of any given series of compounds as determined by analysis, such as the following:

CHLORIDES.

Hydrogen Chloride.

Chlorine	97.24
Hydrogen	2.76
	100.00

Potassium Chloride.

Chlorine	47.51
Potassium	52.49
	100.00

Sodium Chloride.

Chlorine	60.59
Sodium	39.41
	100.00

Silver Chloride.

Chlorine	24.72
Silver	75.28
	100.00

BROMIDES.

Hydrogen Bromide.	
Bromine	98.76
Hydrogen	1.24
	100.00

Potassium Bromide.	
Bromine	67.13
Potassium	32.87
	100.00

Sodium Bromide.	
Bromine	77.62
Sodium	22.38
	100.00

Silver Bromide.	
Bromine	42.56
Silver	57.44
	100.00

IODIDES.

Hydrogen Iodide.	
Iodine	99.21
Hydrogen	0.79
	100.00

Potassium Iodide.	
Iodine	76.41
Potassium	23.59
	100.00

Sodium Iodide.	
Iodine	84.62
Sodium	15.38
	100.00

Silver Iodide.	
Iodine	54.03
Silver	45.97
	100.00

Arranged in this way we do not notice any simple relation existing between the components of this series, except that the quantity of hydrogen is always smaller than that of the chlorine, bromine, or iodine, whilst the quantity of sodium is always smaller than that of potassium, and this again is less than the quantity of silver.

If, however, instead of examining a constant weight of the several compounds we ask how much of the

one constituent in each compound combines with a constant weight of that constituent which is common to several, we shall obtain at once a clear insight into the law of the formation of the compound. In the series of hydrogen compounds, for instance, let us calculate by simple proportion how much chlorine, bromine, and iodine combine with the unit weight of hydrogen. We then obtain for the composition of these compounds:

Hydrogen Chloride.		Hydrogen Bromide.		Hydrogen Iodide.	
Chlorine ..	35.18	Bromine ..	79.36	Iodine ...	125.90
Hydrogen .	1.00	Hydrogen .	1.00	Hydrogen	1.00
	36.18		80.36		126.90

Continuing our calculation, how much of the metals, potassium, sodium, and silver, unite with 35.18 parts by weight of chlorine to form chlorides, with 79.36 parts of bromine to form bromides, and with 125.90 parts of iodine to form iodides. The result is as follows:

CHLORIDES.

Potassium Chloride.		Sodium Chloride.		Silver Chloride.	
Chlorine ...	35.18	Chlorine ...	35.18	Chlorine ...	35.18
Potassium .	38.86	Sodium ...	22.88	Silver ...	107.12
	74.04		58.06		142.30

BROMIDES

Potassium Bromide.		Sodium Bromide.		Silver Bromide.	
Bromine ...	79.36	Bromine ...	79.36	Bromine ...	79.36
Potassium .	38.86	Sodium ...	22.88	Silver ...	107.12
	118.22		102.24		186.48

IODIDES.

Potassium Iodide.		Sodium Iodide.		Silver Iodide.	
Iodine	125.90	Iodine	125.90	Iodine	125.90
Potassium	38.86	Sodium	22.88	Silver	107.12
	164.76		148.78		233.02

A remarkable relation becomes apparent, for it is clear that the same weights of the metals potassium, sodium, and silver, which combine with 35.18 parts of chlorine to form chlorides, also combine with 79.36 parts of bromine to form the bromides, and with 125.90 parts of iodine to form the iodides. In other words, if we replace the 35.18 parts by weight of chlorine in each of these compounds by 79.36 parts of bromine, we get the bromides, and if by 125.90 parts of iodine we obtain the iodides of the metals. Hence one and the same weight of metals, 38.86 of potassium, 22.88 of sodium, and 107.12 of silver, has the power of forming compounds with the precise quantities of chlorine, bromine, and iodine respectively, which unite with 1 part by weight of hydrogen, to form the hydrides of these elements. These quantities of the elements in question are called equivalent quantities, because they are the amounts of them which will combine with the same weight of some other element.

$$\left.\begin{array}{r}38.86\text{ of potassium}\\ 22.88\text{ of sodium}\\ 107.12\text{ of silver}\\ 1.00\text{ of hydrogen}\end{array}\right\}\begin{array}{c}\text{combine}\\ \text{with}\end{array}\left\{\begin{array}{l}35.18\text{ of chlorine}\\ 79.36\text{ of bromine}\\ 125.90\text{ of iodine}\end{array}\right\}\begin{array}{l}\text{respec-}\\ \text{tively.}\end{array}$$

Similar results are obtained from the examination of the compounds of all the other elements, so that a number may be assigned to each element which is

termed the combining weight or equivalent weight of the element.

Taking an example from another group of chemical compounds we find that the well-known oxides of hydrogen, lead, copper, mercury, and cadmium possess the following percentage composition:

OXIDES.

Water.		Lead Oxide.		Copper Oxide.	
Hydrogen	11.19	Lead	92.82	Copper	79.89
Oxygen	88.81	Oxygen	7.18	Oxygen	20.11
	100.00		100.00		100.00

Mercury Oxide.		Cadmium Oxide.	
Mercury	92.63	Cadmium	87.54
Oxygen	7.37	Oxygen	12.46
	100.00		100.00

While the corresponding sulphides exhibit the following composition:

SULPHIDES.

Sulphuretted Hydrogen.		Lead Sulphide.		Copper Sulphide.	
Hydrogen	6.01	Lead	86.58	Copper	66.48
Sulphur	93.99	Sulphur	13.42	Sulphur	33.52
	100.00		100.00		100.00

Mercury Sulphide.		Cadmium Sulphide.	
Mercury	86.18	Cadmium	77.81
Sulphur	13.82	Sulphur	22.19
	100.00		100.00

If, as before, we now compare the quantity of each element united with one and the same weight of oxygen, taking 7.94 parts of this element, because this is the amount of it which combines with one part of hydrogen and is, therefore, the equivalent weight of oxygen with respect to hydrogen, we get the following numbers:

Water.		Lead Oxide.		Copper Oxide.	
Hydrogen	1.00	Lead	102.68	Copper	31.55
Oxygen	7.94	Oxygen	7.94	Oxygen	7.94
	8.94		110.62		39.49

Mercury Oxide.		Cadmium Oxide.	
Mercury	99.25	Cadmium	55.80
Oxygen	7.94	Oxygen	7.94
	107.19		63.74

And, if we investigate the sulphides, we find that one and the same weight of sulphur, 15.91 parts by weight, unites with weights of these elements to form sulphides, which are identical with the amounts that combined with 7.94 parts by weight of oxygen to form oxides, thus:

Sulphuretted Hydrogen.		Lead Sulphide.		Copper Sulphide.	
Hydrogen	1.00	Lead	102.68	Copper	31.55
Sulphur	15.91	Sulphur	15.91	Sulphur	15.91
	16.91		118.59		47.46

Mercury Sulphide.		Cadmium Sulphide.	
Mercury	99.25	Cadmium	55.80
Sulphur	15.91	Sulphur	15.91
	115.16		71.71

Hence we see again that the amounts of these elements which unite with an equivalent of oxygen also combine with an equivalent of sulphur, so that

1.00 part by weight of hydrogen	combines with	7.94 of oxygen
102.68 parts by weight of lead		15.91 of sulphur
31.55 " " " " copper		
99.25 " " " " mercury		
55.80 " " " " cadmium		

and these are, therefore, the equivalent weights of these elements.

When one element combines with another in more than one proportion it is said to have more than one equivalent, and since the amounts of one element which combine with a fixed weight of a second are in a simple ratio to one another, it follows that the several equivalents of an element must also stand in a simple ratio to one another. Iron for example forms several different compounds with oxygen, two of which have the following composition as determined by analysis:

	Ferrous Oxide.	Ferric Oxide.
Iron	77.54	69.97
Oxygen	22.46	30.03
	100.00	100.00

Calculating the amount of iron combined with the equivalent (7.94 parts) of oxygen we find

	(1)	(2)
Iron	27.75	18.50
Oxygen	7.94	7.94
	35.69	26.44

These amounts of iron are, however, in the simple ratio of 2:3, 27.75 being the equivalent of iron in ferrous oxide and 18.50 in ferric oxide.

It will be seen, therefore, that combination always takes place between certain definite and constant proportions of the elements or between multiples of these.

The Atomic Theory becomes truly a chemical atomic theory inasmuch as it supposes the atoms of different elements not to possess the same weights, but to be characterized by different weights. Thus the atom of oxygen is 15.88 times as heavy as the atom of hydrogen, and the weights of the atoms of oxygen and chlorine are as 15.88 to 35.18. Having made these assumptions we can understand why combination always takes place between certain amounts of the elements or between multiples of these amounts, since combination being supposed to take place between some number of atoms of each of the elements which unite, it follows that the amounts which combine must be some finite multiple of the weights of the atoms.

It is thus clear that the atomic theory accounts for the formation of all compounds which are found to exist, but it is equally evident that it in no way decides how many compounds can be formed by any two or more elements. This at present can only be learned by experiment, but we are not without indications that a time approaches when this further problem will receive a theoretical solution.

COMBINING VOLUMES OF GASES.

The relation existing between the volumes of gases when they combine together has been found to be a very simple one, inasmuch as the densities of all the

elements known in the gaseous state are identical with their atomic weights, or, what is the same thing, the atoms in the gaseous state all occupy the same space.

Thus the density and combining weight of oxygen are alike 16 (actually 15.88 with hydrogen as 1), or, oxygen is 16 times heavier than hydrogen, and density and combining weight of nitrogen are alike 14, or, nitrogen is 14 times heavier than hydrogen, the density of chlorine is 35.5, that of sulphur vapor 32, and so on. Remembering this fact, it is easy to calculate the absolute weight of a given volume, say one litre of these different gases, when we know that one litre of hydrogen at the standard pressure and temperature weighs 0.08936 grams. Thus 1 litre of oxygen, under the same circumstances, weighs......$16 \times 0.08936 = 1.430$ grams

1 litre of nitrogen	weighs	$14 \times 0.08936 = 1.251$	"
" chlorine	"	$35.5 \times 0.08936 = 3.172$	"
" sulphur vapor	"	$32 \times 0.08936 = 2.860$	"

With respect to compounds, the density of a compound gas is one-half its molecular weight, or the molecule of a compound gas occupies the space of 2 atoms of hydrogen.

Thus the density of water-gas, or steam, H_2O, is $\frac{18}{2}$ or 9, that is, it is nine times heavier than hydrogen, the density of hydrochloric acid, HCl, is $\frac{36.5}{2}$ or 18.25, that of ammonia, NH_3, $\frac{17}{2}$ or 8.5, that of carbonic acid, CO_2, $\frac{44}{2}$ or 22.

Hence the weights of 1 litre of these compounds, estimated at 0° C and 760 mm., are as follows:

1 litre of steam	weighs	9×0.08936	gram
" ammonia	"	8.5×0.08936	"
" hydrochloric acid	"	18.25×0.08936	"
" carbonic acid	"	22×0.08936	"

The symbol for water, H_2O, therefore, not only indicates that it is composed of 2 parts by weight of hydrogen and 16 of oxygen, but also that 2 volumes of hydrogen have united with 1 volume of oxygen to form 2 volumes or one molecule of water gas. The symbol NH_3 denotes that 3 volumes of hydrogen and 1 volume of nitrogen have united to form 2 volumes, one molecule, of ammonia, whilst the symbol HCl shows that 2 volumes of hydrochloric acid gas contain 1 volume of chlorine and 1 of hydrogen.

We have seen that 28 parts by weight of nitrogen unite with 32 parts of oxygen to form nitrogen dioxide, the density of this compound is, however, found by experiment to be 15, hence its molecular weight is 30, consisting of 14 parts by weight of nitrogen to 16 of oxygen, or 1 volume of each constituent, and its formula must, therefore, be NO.

Nitrogen and oxygen do not readily combine together, but under certain circumstances they are found to do so, thus, if a series of electric sparks are passed through a glass vessel filled with dry air, the presence of red colored vapors, possessing a peculiar acrid smell, is soon noticed. These consist of nitrogen tri- and tetroxides, formed by the union of the nitrogen and oxygen of the air.

From the result of recent experiments the combining weight of oxygen has been taken as 15.88 instead of 16. In the calculations used in this work, the older number 16 has been used in most cases for the sake of simplicity.

CRYSTALLIZATION.

Chemical substances, when they pass from the liquid or gaseous into the solid state, assume some definite geometric form, or are said to crystallize. Crystals are produced when a substance, such as nitre, is dissolved in water and the solution allowed gradually to evaporate, or when a body, such as sulphur, is melted and allowed to solidify by cooling. Crystalline bodies exhibit, in addition to their regular form, a peculiar power of splitting in certain directions more readily than in others, called cleavage, as well as in many cases the property of allowing the rays of light and heat to pass more readily in one direction than another, giving rise to the well-known phenomena of double refraction.

Inorganic bodies which do not exhibit these peculiarities, or assume crystalline structure, are said to be **amorphous**, such as glass and glue. But certain highly complicated structures found in the vegetable and animal world exhibit a structure which, although it is non-crystalline, is not devoid of arrangement, and to which the name organized or cellular structure has been given.

As a rule, every substance possesses a definite form in which it always crystallizes, and by which it can be distinguished. When a crystal is formed from aqueous solution, the smallest visible particle possesses the complete form of the largest crystal, and simply increases in size without undergoing any change of form.

Certain substances exhibiting a similarity in their chemical constitution are found to crystallize in the same forms, these are said to be **isomorphous**; when the same body occurs crystallized in two different systems, it is said to be **dimorphous.**

DETERMINATION OF THE MOLECULAR WEIGHTS OF GASES.

For this purpose it is only necessary to determine the specific gravity of the gas, air being usually taken as the practical unit of comparison. A large glass balloon, capable of holding from 1 to 10 litres of the gas, is employed and is first of all freed from air as far as possible by the vacuum pump, and weighed. In weighing a body of such large volume it is essential to make allowance for the buoyancy of the air, since the body to be weighed appears to be lighter than it really is by an amount equal to the weight of the air which it displaces. This is best accomplished by suspending a vessel of similar size and shape to the other arm of the balance, by which arrangement the effect of buoyancy is neutralized, each of the vessels being affected in the same manner, and at the same time the uncertainties of calculation due to the varying temperature, pressure, and moisture of the atmosphere are avoided, as well as any inaccuracy due to condensation on the surface of the glass. As soon as the weight of the empty vessel has been ascertained, it is removed from the balance and filled with the pure dry gas, the temperature and pressure being carefully observed. The globe is then reweighed with the same precautions as before. One additional correction must be here noticed, as it gives rise to a considerable error especially when light gases are being weighed. The capacity of a vacuous globe of glass is found to be perceptibly less than that of the same globe when filled with gas at

the pressure of the atmosphere, and hence the globe displaces less air in the former condition than in the latter. The difference between the corrected weights of the globe empty and filled with the gas is equal to the weight W of the given volume of gas at the observed temperature and pressure. The same globe is then filled with dry air freed from carbonic acid gas, or else with pure hydrogen, and the weight of an equal volume of the latter A under the same conditions thus ascertained. The specific gravity of the gas compared with air is then equal to $\frac{W}{A}$. Since the air has been found to be 14.39 times as heavy as hydrogen and the molecular weight of a gas is twice its density with respect to hydrogen, it is only necessary to multiply the specific gravity by 14.39×2 to obtain the molecular weight of the gas in question.

FLAME AND THE COMBUSTION OF HYDROCARBONS.

The flames which will be first considered are those formed when a jet of combustible gas or vapor, unmixed with oxygen or inert gases, is allowed to burn in an atmosphere of air or oxygen. The simplest flames of this kind with which we are acquainted are those of

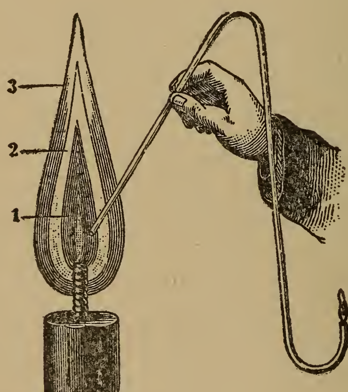

Fig. 1.

hydrogen and carbon monoxide. The flame of either of these gases burning from the end of a tube appears as an incandescent cone, which on examination proves to be hollow, the incandescence only taking place when the gas has mixed with air by diffusion. The hollow nature of these, and indeed of all flames, may be readily shown in various ways. A bent glass tube may be brought into the centre of the flame, when the unburnt

gases will pass up the tube and may be ignited at the other end as in Fig. 1. The head of a match may be thrust quickly into the centre of the flame and held there for some time without the phosphorus catching fire, while the wood will be charred and may even take fire where it is in contact with the hot outer sheath of the cone. A thin platinum wire held horizontally in the flame is seen to glow at two points where it comes in contact with the outer zones in which the combustion is going on, whilst between them it remains cool.

With gases which yield more than one product of combustion the phenomena become more complex. Thus, for example, the flames of cyanogen and sulphuretted hydrogen burning in air are found to consist of two sharply defined cones, possessing different colors. Matters become still more complex in the case of the flames with which we are most familiar, namely, those obtained from a burning candle or from gaseous hydrocarbons. In these flames four distinct regions are usually distinguished; **a** the dark central region, **b** the yellow region, **c** the blue region, **d** the faintly luminous region, which are clearly shown in Fig. 2.

The yellow region, which is, as a rule, the largest, and gives off by far the greatest amount of light, is known in common parlance as the luminous portion.

The separation of carbon particles in a flame of hydrocarbon vapor burning in air cannot be regarded as due to the preferential combustion of the hydrogen, but is mainly owing to thermal changes taking place in the gas under the influence of the heat from the outer sheath of the flame. These changes take place in the dark inner cone, and finally bring about the formation of dense hydrocarbons and carbon particles, which

on reaching the hotter portion of the flame become incandescent, giving the yellow luminous zone.

The thermal changes which thus occur are undoubtedly of a very complex nature, and succeed each other with such rapidity that the experimental determination of their exact nature is a matter of great difficulty. It has been shown by the analysis of the gases taken from a coal-gas flame at different heights that the **unsatu-**

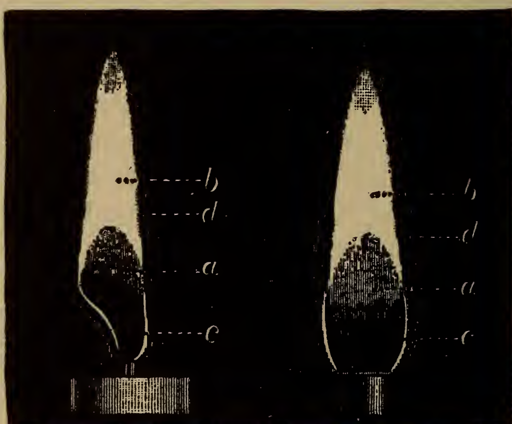

Fig. 2.

rated hydrocarbons decrease very slowly in the dark portion of the flame, but quickly disappear in the luminous zone. The nature of the unsaturated hydrocarbons, however, undergoes a considerable alteration in the non-luminous zone, the amount of acetylene increasing very rapidly and forming 70 per cent of the total unsaturated hydrocarbons when the top of the non-luminous zone is reached. The luminosity of such flames is due in the first place to the formation of acetylene by the action of heat, and the separation of

the carbon is due to the decomposition of the latter gas, which is highly endothermic, and therefore evolves heat in its decomposition, thus further increasing the temperature of the particles at the point of separation. While it is probable that the carbon separation may be in part brought about in this manner, it does not appear likely that this is the sole cause of its production.

The blue portion of the flame is the least in extent of any of the four divisions, and is probably caused by the combustion of hydrocarbons which have become mixed with a sufficient quantity of air to allow them to burn with a scarcely luminous flame, it is probable that the combustion is incomplete, and that the reaction going on here corresponds with that taking place in the inner cone of the flame of the Bunsen burner. The faintly luminous sheath is the region of complete combustion in which those substances which have been incompletely oxidized in the other portions of the flame, chiefly hydrogen and carbonic oxide, are finally converted into water and carbon dioxide. This may be regarded as corresponding with the outer cone of the Bunsen burner flame.

When the hydrocarbon gas or vapor is mixed with other gases before burning, the reactions which take place within the flame are greatly modified, the effect produced depending upon the nature of the added gas. Such gases may consist of incombustible gases such as nitrogen and carbon dioxide, combustible gases such as hydrogen and carbon monoxide, supporters of combustion such as oxygen.

The effect of the addition of either of the first two classes is to reduce the luminosity, and when a sufficient percentage of the dilutent gas is added the flame

becomes non-luminous, the necessary percentage varying according to the gas used.

The loss of luminosity thus brought about is due to the fact that the dilution of the hydrocarbon gas or vapor with other gases raises the temperature at which those thermal decompositions occur leading to the separation of carbon, with the result that, before a sufficiently high temperature for such separation is reached, the products come in contact with the air in the outer flame, and there undergo complete combustion with a non-luminous flame.

The luminosity of a hydrocarbon flame, unmixed with other gases, may also be destroyed if the flame be allowed to impinge on a cold surface, the temperature of the flame being thus reduced to such an extent that in this case also no separation of carbon takes place. This phenomenon may be readily shown by placing a platinum dish containing cold water over a small coal-gas flame, when its luminosity is destroyed, but gradually reappears as the water in the dish becomes hot.

When the hydrocarbon gas or vapor is mixed with a gas which is a supporter of combustion, the flame produced is profoundly modified. When a sufficiently large quantity of oxygen or air is mixed with such gas the flame becomes non-luminous, and at the same time its temperature is greatly increased. Such flames have a distinctly two-coned structure, the inner cone being pale blue, whilst the outer cone is a still paler blue. The most familiar example of this type of flame, which is of the highest technical importance, is seen in the Bunsen gas burner, now so universally used for heating purposes. In this burner the gas emerges from a central jet, Fig. 3, and, passing unburnt up the tube **e e**

Fig. 4, aspirates air with it through the holes **c d**, the mixture then burning with a pale blue or bluish, smokeless flame. If the air holes be closed, the gas burns with the usual luminous flame.

Owing to the presence of oxygen in the gas, the actions which now occur in the inner cone, under the influence of the heat from the outer sheath of flame, are quite different from those taking place in the corresponding portion of a pure hydrocarbon flame. They no longer consist simply of thermal decompositions, but of a direct combustion of the hydrocarbon with a limited supply of oxygen, forming products which are not completely oxidized, but which then burn in the outer cone, the inner cone becomes greener and smaller, and when the mixture becomes sufficiently explosive the flame does not pass down the tube as a whole, but the green inner cone detaches itself from the

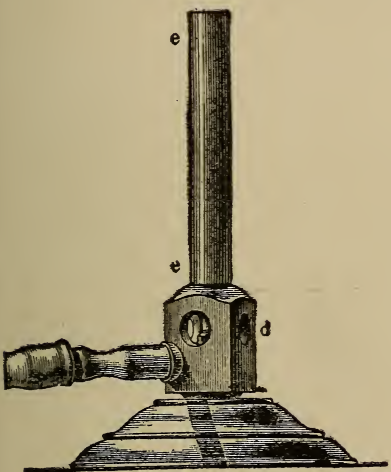

Fig. 3.

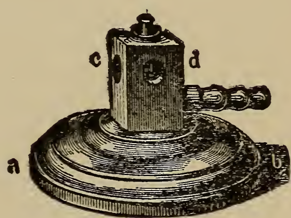

Fig. 4.

outer one and passes down the tube. If the tube be constricted lower down, the descending cone is arrested at that point, the speed of the gases being there greater, and it continues to burn there whilst the outer cone remains in its former position.

LIQUIDS.

It is evident the distinction between gas and vapor is only one of degree, for a vapor is simply a gas below its critical temperature. The same laws according to which the volumes of gases vary under change of temperature and pressure apply also to vapors, at any rate when they are examined at temperatures considerably above their points of condensation. When a vapor is near its point of condensation its density increases more quickly than the pressure, and as soon as the point is reached the least increase of the pressure brings about a condensation of the whole to a liquid.

The essential difference between a liquid and the vapor from which it is produced lies in the fact that the liquid possesses a definite surface tension, and consequently occupies a definite volume limited by a surface, and is not capable, like a gas, of filling any space into which it may be brought. At the surface of the liquid evaporation occurs at all temperatures, and this continues until the pressure of the vapor reaches a certain definite value which depends on the temperature. This value is known as the vapor pressure of the liquid at that particular temperature, and is always reached when an excess of the liquid is present. When the vapor pressure is equal to the superincumbent pressure the liquid is said to boil, the corresponding temperature being termed the boiling point of the liquid. The boiling points usually quoted refer to the normal at-

mospheric pressure of 760 mm. of mercury, and an atmospheric temperature of 0° Centigrade.

Liquids possess a notable vapor pressure below their boiling-points, thus water gives off vapor at all temperatures, and even slowly evaporates when in the solid state, for the pressure of the vapor coming from ice at —10° is 0.208 mm. According to experiments there is a limit beyond which evaporation cannot be detected, thus mercury gives out a perceptible amount of vapor during the summer, but that none can be detected during the winter, and certain compounds which can be volatilized at 150°, undergo no perceptible evaporation when kept for years at the ordinary temperature. This maximum pressure or density of a vapor is not altered by the presence of other gases, or, in other words the quantity of a liquid which will evaporate into a given space is the same whether the space is a vacuum or is filled with another gas. The above conclusions can only be considered as approximately true, as about 2 per cent more vapor ascends into a space filled with gas than into a vacuum, while at considerable but equal distances from the boiling-point the pressures of volatile liquids are by no means equal.

MATTER.

Matter is capable of assuming three different states or conditions: the solid, the liquid, and the gaseous. Of these, the first two have, for obvious reasons, been recognized from the earliest ages, as accompanying very different kinds of substances. It is, however, only within a comparatively short time that we have come to understand that just as there are many distinct kinds of solids and liquids, so there are many distinct kinds of gases. These may be colorless and invisible, but they can readily be shown to differ one from another. A peculiar gas, which we now know as carbonic acid gas, or carbon dioxide, was obtained by the action of dilute acids on marble, to this gas was given the name of **fixed air**, because it is fixed in the alkaline carbonates, which at that time were called the mild alkalies, in contradistinction to the caustic alkalies. This invisible gas does not, like air, support the combustion of a taper, and, unlike air, it renders clear lime-water turbid, it is also much heavier than air, as can be shown by pouring it downwards from one vessel to another, by drawing it out of a vessel by means of a syphon, or by pouring it into a beaker glass previously equipoised at one end of the beam of a balance as in Fig. 5. That the gas has actually been poured out is seen either by a burning taper being extinguished when dipped into the beaker glass, or by adding some clear lime-water, which then turns milky.

The gas which was formerly termed **inflammable air,** and obtained by the action of dilute acids on metallic zinc or iron, is also a peculiar and distinct substance, to which we now give the name of hydrogen gas. It is

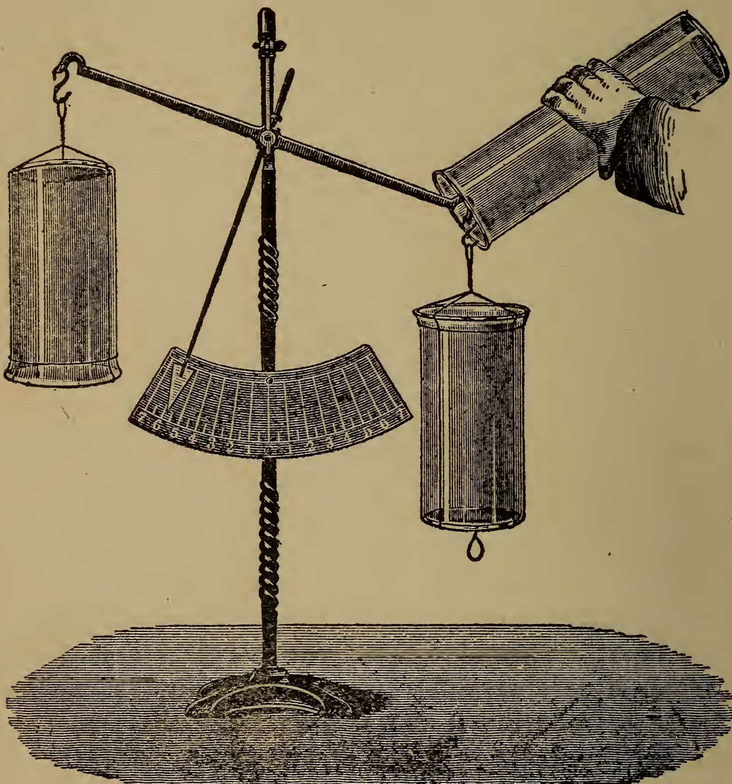

Fig. 5.

so much lighter than air that it may be poured upwards, and takes fire when a light is brought in contact with it, burning with a pale blue flame. Soap-bubbles blown with hydrogen ascend in the air, and if

hydrogen be poured upwards in the equipoised bell-jar hung mouth downwards on the arm of the balance as in Fig. 6, the equilibrium will be disturbed, and the arm with the bell-jar will rise.

Fig. 6.

Chemists have come to the conclusion that matter is indestructible and that in all cases of chemical action in which matter disappears, the loss is apparent only, the solid or liquid being changed into an invisible gas,

the weight of which is, however, exactly identical with that of its component parts. We only require to allow a candle to burn for a few minutes in a clean flask filled with air in order to show that the materials of the candle, hydrogen and carbon, unite with the oxygen of the air to form, in the first place, water, which is seen in small drops bedewing the bright sides of the flask, and in the second, carbon dioxide or carbonic acid gas, whose presence is revealed to us by lime-water being thereby turned milky. The fact that the sum of the weights of the products of combustion (water and carbon dioxide) is greater than the loss of weight sustained by the candle is clearly shown by an experiment made by means of the apparatus shown in Fig. 7, which consists of a tube equipoised on the arm of a balance. In the long vertical tube a taper is placed, the other end of the system being attached to a gasholder filled with water, which, on being allowed to run out, causes a current of air to pass through the tube, and thus maintains the combustion of the taper. The water and carbonic acid gas which are formed are absorbed in the bent tube, which contains caustic potash. After the taper has burnt for a few minutes, the apparatus is disconnected from the gasholder and allowed to vibrate freely, when it will be found to be appreciably heavier than it was before the taper had burnt, the explanation being that the excess of weight is due to the combination of the carbon and hydrogen of the wax with the oxygen of the air.

The statement that matter is indestructible is based entirely upon the evidence of experiment, and many investigations have been carried out to test its validity. These have taken the form of weighing two substances,

such as silver and iodine, mercury and iodine, or iodine and sodium sulphite, as carefully as possible, and then allowing them to unite or react chemically and finally weighing the products. In the experiments the silver and iodine were separately dissolved and allowed to

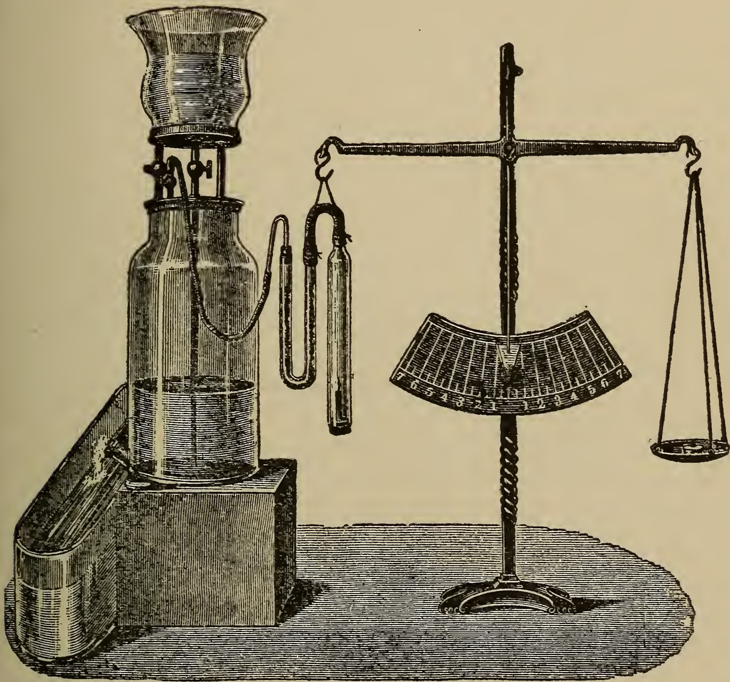

Fig. 7.

react, and the resulting silver iodide was then collected and weighed, but in the other cases the chemical change was allowed to occur within sealed-up vessels, so that no mechanical loss could take place. The result has been that no definite change of weight has in any case been observed. The accuracy of the statement that

matter is indestructible is therefore true within the limits of accurate weighing which have at present been attained.

It is the aim of the chemist to examine the properties of all the different substances which occur in nature, so far as they act upon each other, or can be made to act so as to produce something different from the substances themselves, to ascertain the circumstances under which such chemical changes occur, to discover the laws upon which they are based, and to investigate the relation between the properties of substances and their chemical composition. In thus investigating terrestrial matter it is found that all the various forms of matter with which we are surrounded, or which have been examined, can be divided into two great classes.

Elementary Bodies. Elements, or simple substances, out of which no other two or more essentially differing substances have been obtained.

Compound Bodies, or compounds, out of which two or more essentially differing substances have been obtained.

Of these, and their compounds with each other, the whole mass of our globe, solid, liquid, and gaseous, is composed.

LIST OF ELEMENTS.

	Atomic Weight. H=1.		Atomic Weight. H=1.
Aluminum ....Al	26.9	Neon ....Ne	19.9
Antimony ....Sb	119.3	Nickel ....Ni	58.3
Argon..........................A	39.6	Nitrogen ....N	13.93
Arsenic ....As	74.4	Osmium ....Os	189.6
Barium ....Ba	136.4	Oxygen....................O	15.88
Bismuth...................Bi	206.9	Palladium ....Pd	105.7
Boron ....B	10.9	Phosphorus ....P	30.77
Bromine...................Br	79.36	Platinum ....Pt	193.3
Cadmium ....Cd	111.6	Potassium ....K	38.86
Cæsium ....Cs	131.9	Praseodymium ..Pr	139.4
Calcium ....Ca	39.8	Radium ....Ra	223.3
Carbon C	11.91	Rhodium ....Rh	102.2
Cerium ....Ce	139.2	Rubidium ....Rb	84.8
Chlorine..................Cl	35.18	Ruthenium ....Ru	100.9
Chromium ....Cr	51.7	Samarium...........Sm	148.9
Cobalt ....Co	58.56	Scandium ....Sc	43.8
Copper ....Cu	63.1	Selenium ....Se	78.6
Erbium ....Er	164.8	Silicon ....Si	28.2
Fluorine ....F	18.9	Silver ....Ag	107.12
Gadolinium ....Gd	155.	Sodium ....Na	22.88
Gallium ....Ga	69.5	Strontium ....Sr	86.94
Germanium ....Ge	71.9	Sulphur ....S	31.83
Gold ....Au	195.7	Tantalum ....Ta	181.6
Helium ....He	4.	Tellurium ....Te	126.6
Hydrogen ....H	1.000	Terbium................Tb	158.8
Indium.....................In	131.1	Thallium ....Tl	202.6
Iodine I	125.90	Thorium ....Th	230.8
Iridium ....Ir	191.5	Thulium ....Tm	169.7
Iron......... ....Fe	55.5	Tin ....Sn	118.1
Krypton..................Kr	81.2	Titanium ....Ti	47.7
Lanthanum...........La	137.9	Tungsten ....W	182.6
Lead ....Pb	205.35	Uranium ....U	236.7
Lithium Li	6.98	Vanadium V	50.8
Magnesium ....Mg	24.18	Xenon X	127.
Manganese...........Mn	54.6	Yttrbium.............Yb	171.7
Mercury ....Hg	198.5	Yttrium ....Yt	88.3
Molybdenum.......Mo	95.3	Zinc ....Zn	64.9
Neodymium ....Nd	142.5	Zirconium ....Zr	89.9

PROPERTIES OF GASES.

The chemist has to deal with matter in all its various states, solids, liquids, and gases alike being the objects of his examination. The study of gases in particular has, as we have seen, led to most important results in the theoretical branch of the science, the system of formulæ now employed being in fact founded upon observations made upon matter in the state of gas.

Boyle's Law. The gaseous condition of matter is well defined to be that in which it is capable of indefinite expansion. If a quantity of gas as small as we please be placed in a closed vacuous space, however large, the gas will distribute itself uniformly throughout that space. The relation between the volume and pressure of a gas, the temperature remaining constant, is expressed by the well-known law of Boyle, that the volume of the gas varies inversely as the pressure, from which it follows that the product of pressure and volume remains constant whatever be the pressure. This is expressed by the equation $PV=C$, thus when the pressure is doubled, the volume is halved, the product of the two remaining the same. When a number of gases which do not act chemically on each other are mixed together, the total pressure exerted is equal to the sum of the separate pressures which each gas would exert if it alone occupied the whole space, or as it may be otherwise expressed, the total pressure of a mixture of gases is equal to the sum of the partial pressures of its constituent gases.

Thus, if a litre of oxygen at a pressure of 0.2 of an atmosphere and a litre of nitrogen at a pressure of 0.8 of an atmosphere be mixed and the volume again brought to 1 litre, the total pressure will be 1 atmosphere, this being the sum of the partial pressures of the oxygen and nitrogen.

The adoption of the pressure of a column of mercury 760 mm. high as the normal or standard pressure leads to the anomaly that the mass of a given volume of gas, under standard conditions, varies at different places on the earth's surface, since the pressure exerted by a column of 760 mm. of mercury varies with the latitude and the height above sea level, owing to the variation in the intensity of gravitation. Hence when the weight of a given volume is quoted the locality for which it has been determined must also be stated.

PROPERTIES OF SOLUTIONS.

A remarkable analogy exists between the properties of substances in dilute solution and those of gases. This conclusion has been chiefly derived from the study of the interesting phenomena now to be described. When a solution of a crystalloid substance in water is placed in a vessel closed by a porous membrane, such as a piece of parchment, and the whole immersed in pure water, it is found that the dissolved substance gradually passes outwards through the film of parchment, whilst water passes inwards, until, after a sufficient time has elapsed, equilibrium is established and the liquid has the same composition both inside and outside the membrane, this process being known as osmosis.

Diaphragms of other substances can, however, be obtained which allow the water to pass freely through them in the same way as the parchment, but prevent the outward passage of the dissolved substance, and are therefore said to be "semipermeable." Such a diaphragm can be prepared by filling an ordinary porous cell with a dilute solution of potassium ferro-cyanide and simultaneously immersing it in one of copper sulphate. These two substances gradually diffuse into the porous walls of the cell and produce an insoluble layer of copper ferrocyanide, which is found to be semipermeable for solutions of many salts and other bodies. If now such a cell, containing a dilute solution of sugar, be placed in a vessel containing pure water, the latter

is found to pass inwards through the film, whilst the sugar does not pass out, and consequently the level of the liquid within the cell rises. If, however, the cell be completely filled and connected with an arrangement for measuring the pressure it will be found that the latter gradually increases, but after a time becomes constant. The pressure thus observed is termed the osmotic pressure of the solution, and is found to depend only upon the nature of the dissolved substance, the concentration of the solution, and the temperature.

This osmotic pressure plays the same part in the theory of dilute solutions as the gaseous pressure in that of gases, and is found to follow the same laws. Thus when the strength of the solution is doubled the osmotic pressure becomes twice as great, when it is halved the pressure falls to one-half, and so on. Now doubling the strength of a solution is in reality halving the volume occupied by the unit weight of the dissolved substance, so that the law that the osmotic pressure of a dilute solution varies directly as its concentration corresponds exactly with Boyle's law of gases.

VOLUME OF GASES.

Gases are known as compressible fluids, and liquids as incompressible fluids. Liquids really are compressible, but only to a very slight extent. Like gases, they recover volume on removal of the pressure. The law representing the relation between the volumes of a gas and the pressures to which the gas is subjected is a very simple one. It is termed Boyle's or Mariotte's Law, from the names of the discoverers. It states that the volume occupied by any gas is inversely proportional to the pressure to which it is subjected. Thus the volume 1 under pressure 1 becomes the volume 2 under the pressure $\frac{1}{2}$, the volume 3 under the pressure $\frac{1}{3}$, the volume $\frac{1}{2}$ under the pressure 2, and the volume $\frac{1}{3}$ under the pressure 3, and so on.

The instrument which serves to measure the pressure exerted by the air is termed a barometer. This in its simplest form consists of a straight glass tube, about 800 mm. (33 inches) in length, closed at one end, and furnished with a millimeter scale. This tube is filled with mercury, and the open end placed downwards in a basin containing the same metal. It is then seen that the mercury sinks in the tube to a point about 760 mm. from the surface of the metal in the basin. It is sustained in this position by the pressure of the air. When this pressure increases, the height of the sustained column becomes greater when it diminishes, the level of the mercury in the tube falls. All gases generated at the earth's surface are subject

to this pressure, and their volumes increase or diminish according to the above law, as the superincumbent pressure becomes less or greater. In estimating the volume of hydrogen which can be collected from a given weight of zinc and sulphuric acid, it is clear that we require to know not only the temperature at which the gas is collected, but also the atmospheric pressure under which it is measured, and in order to be able to compare the bulks of two gases, we must always compare them under like conditions of temperature and pressure. For this purpose we compare all the volumes of gases at the standard temperature of 0° C, and under the standard pressure of 760 millimetres of mercury.

DIFFUSION OF GASES.

Another physical property of gases is that of diffusion. Gases which, when mixed together, do not combine chemically, have the power of becoming intimately mixed together, even when differing in specific gravity, and when the heavier gas is placed at the bottom, and both remain at rest. This important property is called the diffusive power of gases. The rate at which gases diffuse varies greatly. Thus, a bottle filled with hydrogen lost 94.5 per cent of this gas when left exposed to the air in the same time as that in which a bottle of carbonic acid lost only 47 per cent of this gas in the same way. Gaseous diffusion goes on through the minute pores of certain solids, such as stucco, or thin plates of graphite; the different diffusive rates of air and hydrogen may be well seen by fixing a thin piece of stucco on to one end of a glass tube open at the other end, and filling this with hydrogen, on plunging the open

end into water a steady rise of this liquid in the tube is noticed, and after some time the whole of the hydrogen is found to have disappeared, and the tube contains only pure air. Experiments made upon this subject have shown that the velocity of diffusion of different gases is inversely proportional to the square roots of their densities, thus 4 volumes of hydrogen will pass through the diaphragm in the same time that 1 volume of oxygen is able to do so, oxygen being approximately sixteen times as heavy as hydrogen. This property of gases has an important bearing upon the atmosphere of towns and dwelling-rooms, which is kept pure to a great extent by this diffusive power of gases.

The following table gives the rates of diffusion of several gases, that of air taken to be equal to 1, compared with the inverse square roots of their densities, air also taken as their unit:

Gas.	Density air=1.	$\frac{1}{\sqrt{\text{density}}}$	Velocity of diffusion air=1.
Hydrogen	0.06926	3.7790	3.830
Nitrogen	0.97130	1.0150	1.014
Oxygen	1.10560	0.9510	0.949
Carbon Dioxide	1.52900	0.8087	0.812

THE ELEMENTS.

ALUMINUM.

Symbol—Al. **Atomic Weight—26.9.**

This metal occurs in large quantities combined with silicon and oxygen in felspar and all the older rocks, and also in clay, marl, slate, and in many crystalline minerals. Metallic aluminum is obtained by passing the vapor of aluminum chloride over metallic sodium. It has recently been manufactured on a large scale and, from its lightness and its bright lustre, it has been used for the metallic portions of optical instruments as well as for ornamental work.

Clay is an aluminum silicate resulting from the disintegration and decomposition of felspar by the action of air and water, the soluble alkali being washed away. Kaolin or porcelain clay is the purest form of disintegrated felspar, containing no iron or other impurities. There are many very beautifully crystalline minerals, consisting of aluminum silicates combined with silicates of the metals of the alkalies and alkaline earths, among others, garnet, idocrase, and mica.

Aluminum salts can be detected when in solution by giving with ammonia a white precipitate, insoluble in excess, but soluble in caustic soda, and by assuming a blue color when moistened with cobalt solution and heated before the blowpipe.

Aluminum and Oxygen Compounds.

Alumina: Formula—Al_2O_3.

Alumina is the only oxide of aluminum known. It occurs native in a nearly pure and crystalline state as corundum, ruby, sapphire, and in a less pure state as emery. Alumina is prepared by adding ammonia to a solution of alum, a white precipitate of the hydroxide falls down, and this on being heated yields a white amorphous powder of pure alumina. This substance is attacked with difficulty by acids, but the hydrate is easily soluble in acids and in the fixed caustic alkalies. Alumina acts as a weak base, the commonest aluminum salts are the alums, and their solutions have an acid reaction. Alumina is largely used in dyeing and calico-printing as a mordant, as it has the power of forming insoluble compounds called lakes with vegetable coloring matter, and thus renders the color permanent by fixing it in the pores of the cloth so that it cannot be washed out: such colors are termed fast.

Aluminum and Chlorine Compounds.

Aluminum Chloride: Formula—Al_2Cl_6.

Aluminum chloride is a volatile white solid body, obtained by heating a mixture of alumina and charcoal in a current of chlorine gas, it is used in the manufacture of the metal.

Aluminum, Sulphur and Oxygen Compounds.

Aluminum Sulphate: Formula—Al_23SO_4.

Aluminum sulphate is a soluble salt prepared on a large scale for the use of the dyer by decomposing clay, by acting upon it with sulphuric acid. The solid

mixture of silica and aluminum sulphate thus obtained goes by the name of alum-cake. The most useful compounds of alumina are, however, the alums, a series of double salts, which aluminum sulphate forms with the alkaline sulphates. A large number of other alums are known, in which the isomorphous sesquioxides of iron, chromium, and manganese are substituted for the alumina in common alum.

ANTIMONY.

Symbol—Sb. **Atomic Weight—119.4.**

Metallic antimony occurs native, but its chief ore is the trisulphide. The metal is easily reduced by heating the sulphide with about half its weight of metallic iron, when ferrous sulphide and metallic antimony are formed. Antimony may also be reduced by mixing the ore with coal and heating in a reverberatory furnace. Antimony is a bright bluish-white colored metal, it is very brittle, and can be powdered in a mortar, it melts at 450°, and may be distilled at a white heat in an atmosphere of hydrogen. Antimony undergoes no alteration in the air at ordinary temperatures, but rapidly oxidizes if exposed to air when melted, and, if heated more strongly, it takes fire and burns with a white flame, giving off dense white fumes of antimony trioxide. Antimony is not attacked either by dilute hydrochloric or sulphuric acids. Nitric acid attacks the metal, converting it into white insoluble antimony pentoxide. Nitro-hydrochloric acid dissolves antimony easily. The alloys of antimony are largely used in the arts. Of these, type metal, an alloy of lead and antimony, is the most important, it contains 17 to 20 per cent of the latter metal.

Antimony and Oxygen Compounds.

Antimony Trioxide, Sb_2O_3 and Antimony Pentoxide, Sb_2O_5.

This oxide gives rise to the important series of salts of antimony used in medicine, it is obtained in crystalline needles, which are isomorphous with the rare form of arsenic trioxide. Antimony trioxide has also been observed to crystallize in octahedra, hence these two oxides are said to be iso-dimorphous. The best mode of preparing the pure oxide is by decomposing antimony trichloride with an alkaline carbonate, when the oxide is precipitated as a white powder.

Antimony trioxide dissolves, when boiled with a solution of cream of tartar (hydrogen potassium tartrate), and on concentration the solution deposits crystals of tartar emetic (potassium antimony tartrate), antimony trioxide also dissolves in hydrochloric acid, yielding a solution of the trichloride, which is rendered turbid by addition of water, owing to the formation of an insoluble antimony oxychloride.

Antimony Pentoxide: Formula—Sb_2O_5.

Antimony Pentoxide, sometimes called Antimonic Acid, obtained by acting on antimony with strong nitric acid, or by decomposing the pentachloride of antimony with water. It is a light straw-colored powder, which loses oxygen at a red heat, and is converted into the intermediate oxide $Sb_2O_3Sb_2O_3$. Antimony pentoxide forms salts with the alkalies called antimoniates, from which antimonic acid, $Sb_2O_6H_2$, can be separated as a white powder. The oxides prepared by the two methods above given are found to possess different properties as regards their power of uniting

with bases. That prepared with nitric acid yields monobasic salts, while that obtained from the pentachloride yields dibasic salts. To the first class of salts the name antimoniates, and to the second that of metantimoniates, has been given.

The grey intermediate tetroxide, Sb_2O_4, is obtained by heating the metal in the air until no further change occurs.

Finely-powdered metallic antimony takes fire spontaneously when thrown into chlorine gas, with formation of the chlorides.

ARSENIC.

Symbol—As. **Atomic Weight—74.4.**

Arsenic closely resembles phosphorus in its chemical properties and in those of its compounds, although in physical characters, such as specific gravity and lustre, it bears a greater analogy to the metals. It may be considered the connecting link between these two divisions of the elements, antimony and bismuth being closely connected with it on the one hand, and phosphorus and nitrogen on the other. Arsenic is sometimes found in the free state, but more frequently combined with iron, nickel, cobalt, and sulphur. It is also contained in very small quantities in many mineral springs. In order to separate arsenic from any of the metallic ores in which it occurs, the ore is roasted, or exposed to a current of heated air in a reverberatory furnace, the arsenic combines with the atmospheric oxygen, forming arsenic trioxide, As_2O_3, which is carried in the state of vapor from the furnace into long chambers or flues, in which the trioxide, commonly

known as arsenious acid, or white arsenic, is deposited. Metallic arsenic may be prepared from this oxide by mixing it with charcoal and sodium carbonate, and heating in a closed crucible, the upper part of which is kept cool, arsenic condenses in the cool part of this apparatus as a solid with a brilliant greyish lustre. It tarnishes in the air from oxidation, it has a specific gravity of 5.7 to 5.9, and when heated to dull redness, it volatilizes as a colorless vapor without undergoing fusion, and this vapor possesses a remarkable garlic-like smell. Arsenic when heated in the air takes fire, and burns with a bluish flame, forming arsenic trioxide As_2O_3, when thrown into chlorine, it instantly takes fire, forming arsenic trichloride, $AsCl_3$.

In order to purify the commercial arsenic it is sublimed with the addition of a small quantity of powdered charcoal. On a small scale, arsenic may be purified by introducing the mixture into a glass flask, which is then placed in a large crucible, surrounded by sand, and heated to redness. As soon as the sublimation begins, a loosely-fitting stopper of chalk is placed in the neck of the flask, a second crucible is placed over the first, and the whole heated until the arsenic is sublimed into the upper portion of the flask. In this way crystals of arsenic are obtained, which have a bright metallic lustre.

Arsenic has a steel gray color, and is a good conductor of electricity. If pure arsenic is quickly sublimed in a stream of hydrogen gas it is deposited in the neighborhood of the heated portion of the tube in crystals, but, at a little distance, as a black glittering mass, and, still further on, as a yellow powder. Both the black and the yellow modifications are crystalline. Yel-

low arsenic is best prepared by distilling arsenic in a stream of carbon dioxide, the vapors being passed into a **U**-tube, where the arsenic is condensed by coming into contact with another stream of carbon dioxide which has been cooled, the arsenic is then dissolved in carbon bisulphide, from which it is deposited by evaporation or by cooling to — 70°.

This form of arsenic is extremely sensitive to light, changing quickly into the ordinary one, but a solution in carbon bisulphide may be kept for some time unchanged, whilst both the yellow and the black modifications give metallic arsenic on heating.

A reddish-brown crystalline modification, which is not changed by light, is stated to be deposited from a solution of the yellow arsenic in carbon bisulphide on long standing.

It was formerly supposed that arsenic could not be melted, for when heated under ordinary circumstances to about 450° it passes at once from the solid to the gaseous state. Its melting point lies between those of antimony and silver.

Arsenic and Oxygen Compounds.

Arsenious Oxide: Formula—As_4O_6.
Molecular Weight—361.1.

Arsenious oxide has long been known under the names of white arsenic and arsenious acid. The name of smelting-furnace-smoke is given to this substance because it is obtained by roasting arsenical pyrites, and is emitted during the process in the form of a white smoke which condenses to a white powder.

Arsenious oxide is prepared on a large scale in many metallurgical processes by the roasting of arsenical

ores. The vapors of the oxide which are given off are condensed in long passages or chambers called poison-chambers, formerly in towers termed poison-towers, in the form of crude flowers of arsenic or poison-flower. For the preparation of white arsenic, arsenical pyrites is usually employed. It is obtained as a by-product in the roasting of cobalt ores, which are employed in the manufacture of smalt. When kept for any length of time it becomes opaque, being changed into a porcelain-like mass. This change is due to the passage from the amorphous or vitreous to the crystalline condition, the change commences at the outside of the mass, and gradually penetrates into the interior.

Arsenious oxide is slightly soluble in water. It is very much more soluble in hydrochloric acid than in water, and it may be easily obtained from the solution in the form of large crystals. It also occurs in this form as arsenic-bloom, being found together with native arsenic, having been formed by the oxidation of this substance.

Arsenious oxide serves for the preparation of a large number of other arsenic compounds, especially of the acids of arsenic and their salts. It is also employed in the manufacture of arsenical pigments and is largely used in the manufacture of glass.

Arsenic and Sulphur Compounds.

Arsenic subsulphide, As_3S, Arsenic disulphide, As_2S_2 and Arsenic trisulphide, As_2S_3.

Arsenic subsulphide: Formula—As_3S.

This compound is obtained by adding phosphorus trichloride to an aqueous solution of sodium arsenate,

allowing the mixture to cool, saturating with sulphur dioxide, and allowing the whole to stand for two or three days, when a dark brown precipitate of arsenic subsulphide separates out.

This sulphide, on heating, decomposes, giving realgar which sublimes, and arsenic which remains behind.

It is soluble in yellow ammonium sulphide, forming a solution from which arsenic trisulphide is precipitated on acidification.

Arsenic disulphide or realgar: Formula—As_2S_2.

This compound occurs native as realgar, crystallizing in oblique prisms. These possess an orange-yellow color and resinous lustre, and are more or less translucent, the streak varies from orange-yellow to red. It occurs together with silver and lead ores at Andreasberg in the Harz and other localities, and embedded in dolomite on the St. Gothard, and has been found in minute crystals in Vesuvian lavas.

When it is heated to 150° with a solution of sodium bicarbonate it dissolves in the liquid and afterwards separates out in the crystalline form.

The red arsenic glass or ruby sulphur which occurs in commerce is an artificial disulphide of arsenic prepared in various arsenic works.

Ruby sulphur is a red glassy mass, translucent at the edges. It does not possess a constant composition. The material manufactured at Freiberg contains generally 75 per cent of arsenic and 25 per cent of sulphur, and that made at Reichenstein in Silesia is a mixture of 95 parts of disulphide with 5 parts of sulphur. This body was formerly much used as a pigment, and is still employed in the manufacture of the so-called Indian or white-fire, which is a mixture of

two parts of the disulphide with twenty-four parts of nitre, and burns with a splendid white light when ignited. The disulphide is also employed in tanning, being mixed with lime and employed for removing the hair from the skins.

Arsenic Trisulphide: Formula—As_2S_3.

This substance occurs in nature and is known under the name of orpiment or the yellow sulphide of arsenic. It crystallizes in translucent lemon-colored prisms.

When sulphuretted hydrogen is passed through an aqueous solution of arsenious oxide the liquid becomes of a yellow color, but no precipitate is formed, the liquid containing arsenic trisulphide in the colloidal form. If a small quantity of hydrochloric acid be present, a beautiful yellow precipitate of arsenic trisulphide is at once thrown down. On heating this substance it melts to a yellowish-red liquid which volatilizes without decomposition at a temperature of about 700°. Heated in the air it takes fire and burns with a pale blue colored flame to arsenious oxide and sulphur dioxide.

It dissolves in solutions of the alkali hydroxides, a brown precipitate of arsenic, containing a little sulphur, separating out slowly, whilst the solution contains thio-arsenate and thio-oxyarsenates.

The sulphide of arsenic occurring in commerce is prepared by subliming a mixture of seven parts of pulverized arsenious oxide with one part of sulphur, and it is really a mixture of arsenious oxide with more or less sulphide of arsenic. The material thus prepared, which is very poisonous, from the excess of

arsenious oxide which it contains, was formerly much used as a pigment under the name of King's yellow, but it is now almost entirely superseded by the comparatively innocuous chrome yellow. The yellow sulphide of arsenic is also used in the arts and manufactures, for instance in the printing of indigo colors, and a mixture of orpiment, water, and slaked lime is used in the East under the name of Rusma as a depilatory, its action depending upon the formation of a hydrosulphide of calcium.

BARIUM.

Symbol—Ba. **Atomic Weight—136.41.**

Barium compounds occur somewhat more widely dispersed than those of strontium, the two most common barium minerals being the sulphate, or heavy spar, and the carbonate. The metal barium has not yet been obtained in the coherent state.

Barium and Chlorine Compounds.

Chloride of Barium: Formula—$BaCl_2$.

Barium Chloride is one of the most important compounds of barium, it crystallizes in flat scales containing two atoms of water. It may be prepared by dissolving the native carbonate in hydrochloric acid, and it is largely used as a precipitant for sulphuric acid.

Sulphate of Barium: Formula—$BaSO_4$.

Barium Sulphate occurs native and crystalline as heavy spar. It is one of the most insoluble salts known, and falls as a white crystalline precipitate when any soluble barium salt is brought into a solution of a sulphate. It is used as a paint, and the precipitated

salt is termed blanc fixe, whilst the native heavy spar, when ground, is largely used to adulterate white lead.

Barium and Oxygen Compounds.

Barium Monoxide or Baryta, BaO and Barium Dioxide, BaO_2.

Barium Monoxide: Formula—BaO.

The best way of forming this oxide is to decompose the nitrate by heat. It is a greyish porous mass, which fuses at a high temperature, and takes up water with evolution of much heat, forming a crystalline hydrate. This hydrate is soluble in twenty parts of cold water, and the solution on exposure to the air rapidly absorbs carbonic acid, and becomes milky.

Barium Dioxide: Formula—BaO_2.

When baryta is gently heated in a current of oxygen gas, the two substances combine together to form a dioxide containing twice as much oxygen as baryta, this additional atom of oxygen is, however, evolved at a higher temperature, and it has been proposed to use this decomposition for the manufacture of oxygen from the air. For this purpose, as soon as the dioxide BaO_2 has been reduced to BaO, the temperature is lowered, and air passed over the baryta, this again takes up oxygen, passing into BaO_2, which again is decomposed by a higher temperature. This interesting process has, however, been found not to work in practice. There are no salts known corresponding to this oxide.

BISMUTH.

Symbol—Bi. **Atomic Weight—207.0.**

This metal is found in small quantities in the native state, but occurs more often as a sulphide, it is easily

reduced to the metallic state, and then exhibits a pinkish-white color. Bismuth does not oxidize in dry air at the ordinary temperature, but if heated strongly it burns with a blue flame, forming an oxide, it also takes fire when thrown into chlorine gas. Bismuth dissolves easily in nitric acid. The metal is chiefly used as an ingredient of fusible metal, its compounds are also used in medicine and as pigments.

Bismuth and Oxygen Compounds.

Two oxides of bismuth are known, Bismuth trioxide, Bi_2O_3, and Bismuth pentoxide, Bi_2O_5. The first of these is a pale yellow powder, formed when the metal is roasted in the air, the second oxide is obtained by dissolving the first in potash, and precipitating the pentoxide by nitric acid and heating, it is a reddish-brown powder. Like the corresponding antimony compound, bismuth pentoxide forms with the alkalies soluble salts.

Bismuth and Nitrogen Compounds.

Bismuth nitrate, $Bi(NO_3)_3+5H_2O$, is the most important soluble salt of bismuth, the sulphide, Bi_2S_3, is a black insoluble compound, the trichloride, $BiCl_3$, is obtained by heating the metal in chlorine. One of the most striking peculiarities of the bismuth compounds is, that solutions of the salts become milky on the addition of water, owing to the formation of insoluble basic compounds. Thus $Bi(OH)_2NO_3$ is formed as a white powder, used in medicine, by adding water to a solution of the normal nitrate, and an oxichloride $BiOCl$ is precipitated by adding water to the trichloride. Metallic bismuth is easily reduced from its compounds, before the blowpipe, as a brittle bead.

BORON.

Symbol—B. **Atomic Weight—10.91.**

Boron combined with oxygen and sodium is found as borax in nature, it is also found combined with oxygen alone as boron trioxide. It exists in two forms, crystalline and amorphous. Boron is easily obtained as a grey amorphous powder, by heating fused boron trioxide, B_2O_3, with sodium. Crystallized boron is prepared by heating the amorphous form strongly with aluminium, this metal in the fused state having the property of dissolving boron, which separates out in nearly colorless crystals when the metal cools, just as the graphitoidal form of carbon does from its solution in iron on cooling. In one specimen of these colorless crystals which was analyzed some quantity of carbon was found to be present, hence carbon may be said to have been prepared artificially in the diamond modification. Boron burns when strongly heated in oxygen or in chlorine, forming the oxide or chloride. It is remarkable as being one of the few elements which unite directly with nitrogen.

Amorphous Boron. This modification is obtained by heating boron trioxide, obtained by the ignition of boric acid, with potassium in an iron tube. It may also be obtained by mixing ten parts of coarsely powdered boron trioxide with six parts of sodium, bringing the mixture into a crucible already heated to redness, and covering it with a layer of powdered sodium chloride previously well dried. As soon as the reaction, which is very violent, has subsided, the mass is stirred with an iron rod until all the sodium has been oxidized, and then carefully poured into water acidified with hy-

drochloric acid. The soluble salts dissolve in the water, whilst the boron remains behind as an insoluble brown powder. This is then collected on a filter, and it must be very carefully dried, as it is easily oxidized and may take fire.

The boron obtained in this way is impure, and a similar product is got when boron trioxide or borax is heated with magnesium powder. In this case a boride of magnesium is probably formed which is decomposed by acids. The product left after this treatment, however, still contains about 5 per cent of magnesium and about 1 per cent of hydrogen, which is also found in boron prepared by means of sodium, and is probably present in the form of a solid hydride.

In order to prepare pure boron, 70 grams of magnesium powder are heated with 210 grams of boron trioxide. The product, which consists of boron, accompanied by magnesium boride and borate, is treated with dilute acid which dissolves the borate and the greater part of the boride. In order to remove the last portions of the boride the residue is then fused with borax and again treated with hydrochloric acid, which leaves the boron containing only traces of silicon, iron, and magnesium. The presence of nitride of boron in the final product can only be avoided by carrying out the heating in an atmosphere of hydrogen or by placing the mixture in a crucible lined with titanic acid.

Amorphous boron is a chestnut-brown powder. It is infusible even at the temperature of the electric arc, but volatilizes, the extremities of the electrodes being converted into boron carbide, while if it be heated to 700° in air it burns. It is a more powerful reducing

agent than carbon or silicon, inasmuch as it is oxidized by carbonic oxide and by silica. It combines with bromine at 700°, but not with iodine, and forms compounds with many metals on heating, among which are silver and platinum, while on fusing with cast-iron in a current of hydrogen it replaces the carbon giving a white, hard cast-iron, and if fused with reduced iron produces brown steel. It is acted on by the oxy-acids, is oxidized by water vapor, and combines with nitrogen at a high temperature. It is a non-conductor of electricity, and when freshly prepared and not strongly ignited is slightly soluble in water, imparting to it a yellow color and being precipitated unchanged from its aqueous solution on the addition of acids or salts.

Crystalline or Adamantine Boron. This substance was first obtained in the year 1856. It can be prepared by several processes. Thus, if amorphous boron be pressed down tightly in a crucible, a hole bored in the center of the pressed mass, a rod of aluminium dropped into the hole, and the crucible then heated to whiteness, the boron dissolves in the molten aluminium and separates out in the crystalline form when the metal cools. The aluminium is then dissolved in caustic soda, and thus the insoluble boron is left in large transparent yellow or brownish-yellow crystals. The same modification may be obtained in smaller crystals, which are often joined together in the form of long prismatic needles, by melting together boron trioxide and aluminium. In order to prevent the action of the oxygen of the air upon the fused mass, the crucible in which the operation is conducted must be placed inside a larger one and the space between them filled up with powdered charcoal. In this process, however, the boron

takes up carbon to the amount of from 2 to 4 per cent. This carbon must be in the form of diamond carbon, inasmuch as the boron crystals containing this impurity are transparent, and more transparent the larger the percentage of carbon. In addition to carbon the boron thus prepared is found to contain a certain quantity of iron and silicon from the crucible used. These impurities can be removed by treatment with hydrochloric acid, and afterwards with a mixture of nitric and hydrofluoric acids. The crystals of adamantine boron contain aluminium as well as carbon, and possess a constant composition.

It may also be prepared by heating boric acid with aluminium turnings and sulphur, and treating the cooled mass with water, when aluminium sulphide is decomposed and crystalline boron left.

Boron crystallizes in monoclinic pyramids or prisms, which have a lustre and a hardness exceeded only by that of the diamond, as they scratch both ruby and corundum. When heated in the air or in oxygen it ignites at the same temperature as does the diamond, and then does not oxidize throughout the mass, but becomes covered with a coating of the melted trioxide. Concentrated nitric acid exerts no action upon it, and even aqua regia attacks it but slowly. Boiling caustic soda solution likewise does not act upon it, but if it is fused with the solid alkali it dissolves slowly with formation of sodium borate and with evolution of hydrogen.

Boron and Carbon Compounds.

Boron Carbide: Formula—B_2C_2.

This is prepared by heating a mixture of boric anhydride and carbon in an electric furnace, and forms

a graphite-like powder which melts at a very high temperature, burns with difficulty in oxygen, is insoluble in the usual solvents, but is decomposed by fusion with alkalies.

Another boride of carbon, having the composition CB_6, is prepared by heating sixty-six parts of amorphous boron and twelve parts of carbon, obtained from sugar, in an electric furnace, or by dissolving the two elements in iron, copper, or silver at a very high temperature, and removing the metal by treatment with aqua regia. It forms very hard black lustrous crystals, having a density of 2.41, which are attacked by chlorine below 1,000°, and by oxygen very slowly at that temperature, like the previous compound, it is decomposed by fusion with alkalies. The powder is sufficiently hard to cut the diamond but more slowly than diamond dust.

BROMINE.

Symbol—Br. Atomic Weight—79.36.

This element, which closely resembles chlorine in its properties and compounds, was discovered in the salts obtained by the evaporation of sea-water. It does not occur free in nature, and is, like chlorine, found combined with sodium and magnesium in the waters of certain mineral springs. In order to obtain pure bromine, use is made of the fact that free chlorine liberates bromine from its combinations with metals, forming a metallic chloride. The bromine thus set free may be separated by shaking the liquid up with ether, which dissolves the bromine, forming a bright red solution. On adding caustic potash to this ethereal solution, the color at once disappears, the bromine be-

comes combined, forming the bromide and bromate of potassium, on evaporation of the ether these salts remain, and after ignition to decompose the bromate, the bromide can again be liberated by the action of sulphuric acid and manganese dioxide, exactly as in the case of chlorine.

One part of bromine dissolves in about 30 parts of water at 15°, and this solution possesses bleaching powers, feebler, however, in action than those of chlorine. This bleaching action is caused by the oxidation of the coloring matter, the bromine combining with the hydrogen of the water to form an acid called hydrobromic acid, corresponding in mode of formation and properties to hydrochloric acid.

In order to detect bromine in a mineral water, or to prepare it in small quantities, the following method is employed. The mother-liquor remaining after the brine has been well crystallized is treated with a stream of chlorine gas, so long as the yellow color of the liquid continues to increase in depth. Chlorine has the power of liberating bromine from bromides, itself uniting with the metal, and the bromine being set free, thus:

$$MgBr_2+Cl_2=MgCl_2+Br_2.$$

The addition of excess of chlorine is to be avoided, as a compound of chlorine and bromine is then formed. The yellow liquid is then well shaken with chloroform, which dissolves the bromine, forming, on standing, a brown solution below the aqueous liquid. On adding caustic potash to this solution the color at once disappears, the bromine combining to form the bromide, KBr, and bromate of potassium, $KBrO_3$, thus:

$$3Br_2+6KHO=KBrO_3+5KBr+3H_2O.$$

On further concentrating the solution a mixture of these salts remains, and from these the bromine is again liberated by distilling the liquid with black oxide of manganese and sulphuric acid in a tubulated retort. The decomposition which here occurs is similar to that which takes place in the preparation of chlorine. Dark red fumes of bromine are liberated, and a black liquid condenses in the well-cooled receiver.

If the bromine is required to be anhydrous it must be re-distilled over concentrated sulphuric acid, and if iodine is present this must be got rid of previously by precipitation as subiodide of copper.

By far the greater quantity of the bromine brought into commerce is now manufactured at Stassfurt from the mother-liquor remaining after the separation of the potassium salts contained in the salt deposits, the process adopted consisting in the treatment of the liquors with chlorine under suitable conditions.

Bromine is a heavy mobile liquid, so dark as to be opaque except in thin layers. It is the only liquid element at the ordinary temperature except mercury. It freezes at —7° to a dark brown solid, evaporates quickly in the air, boils at 59°, and crystallizes from carbon bisulphide at —90° in slender dark carmine-red prisms. Bromine possesses a very strong, unpleasant smell, the vapors when inhaled produce great irritation, and affect the eyes very painfully. When swallowed it acts as an irritant poison, and when dropped on the skin it produces a corrosive sore which is very difficult to heal.

In its general properties, as well as in those of its compounds, bromine closely resembles chlorine, al-

though they are not so strongly marked. Thus it bleaches organic coloring matters, but much less quickly than chlorine, and it combines directly with metals to form bromides, though its action is less energetic than that of chlorine. It does not combine at all at ordinary temperatures with metallic sodium, indeed these two substances may be heated together to 200° before any perceptible action commences, whereas bromine and potassium cannot be brought together without combination occurring, sometimes with almost explosive violence. The addition, however, of a drop of water to bromine and clear sodium sets up a lively reaction.

The solution of bromine in water has an orange-red color, it soon loses bromine in contact with the air, and bleaches organic coloring matter. Bromine water is permanent in the dark, but on exposure to sunlight it becomes acid from the formation of hydrobromic acid and evolution of oxygen. Bromine also dissolves readily in chloroform, carbon bisulphide, alcohol, ether, and acetic acid.

Bromine and Hydrogen Compounds.
Hydrobromic Acid: Formula—HBr.

Hydrobromic acid, or hydrogen bromide, is prepared by the action of acids (phosphoric acid) on the bromides, or better by bringing bromine and phosphorus in contact with water, when a violent action occurs, hydrobromic acid and phosphoric acid being formed thus:

$$P+5Br+4H_2O=5HBr+H_3PO_4.$$

It is colorless gas, having a strong acid reaction, and fumes strongly in moist air, it is very soluble in water. When concentrated, the aqueous acid boils (under 760

mm. pressure) at 126°, and contains 47.8 per cent of HBr. Two volumes of this gas contain one of bromine united with one of hydrogen. The aqueous acid neutralizes bases, forming the bromides and water. The gas liquefies at —73°.

CADMIUM.

Symbol—Cd. **Atomic Weight—111.3.**

This is a comparatively rare metal, occurring in small quantities in most zinc ores. In its chemical relations it closely resembles zinc. It is, however, more volatile than the latter metal, and therefore distills over first in the preparation of zinc. Cadmium is a white ductile metal, melting at 315°, it may be easily distinguished and separated from zinc by yielding a bright yellow sulphide which is insoluble in hydrochloric acid.

CALCIUM.

Symbol—Ca. **Atomic Weight—39.7.**

Calcium forms a considerable portion of the plutonic rocks of which the earth is composed, and occurs in very large quantities, forming whole mountain-chains of limestone, chalk, gypsum, and mountain limestone. The metal calcium is obtained by the decomposition of the chloride by the electric current, or by heating the iodide with sodium, it is a light yellow metal which easily oxidizes in the air, and when heated in air it burns with a bright light, lime, the only oxide of calcium, being formed.

Calcium and Chlorine Compounds.

Calcium Chloride: Formula—$CaCl_2$.

This soluble salt is formed when limestone or marble is dissolved in hydrochloric acid. If the solution

be then evaporated, colorless needle-shaped crystals of the hydrated chloride are formed. When these are dried, the substance still retains $2H_2O$, and forms a porous mass which takes up moisture with great avidity, and is much used for drying gases. When this mass is more strongly heated, it fuses and parts with all its water.

Calcium and Oxygen Compounds.

Calcium Oxide or Lime: Formula—CaO.

Pure lime is obtained by heating white or black marble to redness in a vessel exposed to the air. Lime is prepared on a large scale for building and other purposes, by heating limestone in kilns by means of coal mixed with the stone, the carbonic acid escapes, and quick or caustic-lime remains. Pure lime is a white infusible substance, which combines with water very readily, giving off great heat, and falling to a white powder called calcium hydroxide, or slaked lime. The hydrate is slightly soluble in water, 1 part of it dissolving in 730 parts of cold, but only in 1300 parts of boiling water, and forming lime-water, which, like the hydrate, has a great power of absorbing carbonic acid from the air. It is indeed owing to this property that the hardening or setting of mortars and cements made from lime is due. Mortar consists of a mixture of slaked lime and sand. A gradual combination of the lime with the silica occurs, and this helps to harden the mixture. Hydraulic cements, which harden under water, are prepared by carefully heating an impure lime containing clay and silica. A compound silicate of lime and alumina appears to be formed on moistening the powder, which then solidifies, and is unacted

upon by water. Lime is largely used in agriculture, its action being to destroy the excess of vegetable matter contained in the soil, and to liberate the potash for the use of the plants from heavy clay soils by decomposing the silicate.

Calcium, Carbon and Oxygen Compounds.

Calcium Carbonate: Formula—$CaCO_3$.

This salt occurs most widely diffused, as chalk, limestone, coral, and marble, many of those enormous deposits being made up of the microscopic remains of minute sea-animals. Calcium carbonate exists crystalline as calc-spar, or Iceland spar, and also in a different form, so that this substance is dimorphous. The carbonate is almost insoluble in pure water, but readily dissolves when the water contains carbonic acid, giving rise to what is termed temporarily hard water. Such a water deposits a crust of calcium carbonate on boiling, owing to the escape of the carbonic acid. The well-known evil of boiler incrustation is caused by these deposits. The formation of such a crust may be checked, if not avoided, by adding a small quantity of sal-ammoniac to the water, soluble calcium chloride and volatile ammonium carbonate being formed. Water hard with dissolved carbonate may be softened by the addition of lime suspended in water in such quantity that the excess of carbonic acid is neutralized.

Calcium, Chlorine and Oxygen Compounds.

Chloride of Lime: Formula—$CaCl_2$. $Ca2ClO$.

Bleaching Powder, or Chloride of Lime, is a mixture of calcium chloride and calcium hypochlorite, and is

obtained by the action of chlorine upon slaked lime. If a clear solution of bleaching powder is heated with a small quantity of oxide of cobalt or of copper, the oxygen of the hypochlorite is gradually evolved, and calcium chloride left behind. This decomposition depends upon the fact that higher oxides of the metal are at first formed, but these decompose under the influence of heat, and give off oxygen, regenerating the lower oxide, which again attacks another portion of hypochlorite, and thus the process becomes continuous. It is not improbable that the action of manganese dioxide in facilitating the evolution of oxygen from potassium chlorate may depend upon a similar action.

Calcium Sulphate: Formula—$CaSO_4$.

This occurs in nature as a mineral termed Anhydrite, and combined with $2H_2O$ as selenite, gypsum, or alabaster. It is soluble in 400 parts of water, and is a very common impurity in spring water, giving rise to what is termed permanent hardness, as it cannot be removed by boiling. Gypsum when moderately heated loses its water, and is then called plaster of Paris, this when moistened takes up two atoms of water again and sets to a solid mass, and is therefore much used for making casts and moulds.

CARBON.

Carbon: Symbol—C. Atomic Weight—16.

Carbon is remarkable as existing in three distinct forms, which, in outward appearance or physical properties, have nothing in common, whilst their chemical relations are identical. These three allotropic forms

of carbon are Diamond, Graphite or Plumbago and Charcoal. These substances differ in hardness, color, specific gravity, etc., but they each yield on combustion in the air or oxygen the same weight of the same substance, carbonic acid, or carbon dioxide, 12 parts by weight of each of these forms of carbon yielding 44 parts by weight of carbon dioxide. Carbon is the element which is specially characteristic of animal and vegetable life, as every organized structure, from the simplest to the most complicated, contains carbon. If carbon were not present on the earth, no single vegetable or animal body such as we know could exist. In addition to the carbon which is found free in these three forms, and that contained combined with hydrogen and oxygen in the bodies of plants and animals, it exists combined with oxygen as free carbon dioxide in the air, and with calcium and oxygen as calcium carbonate in limestone, chalk, marble, corals and shells. The fact has already been noticed that plants are able, when exposed to sunlight, to decompose the carbon dioxide in the air, liberating the oxygen, and taking the carbon for the formation of their vegetable structure, whilst all animals, living directly or indirectly upon vegetables, absorb oxygen, and evolve carbon dioxide. Thus the sun's rays, through the medium of plants, effect deoxidation or reduction, while animals act as oxidizing agents with respect to carbon.

The element carbon not only combines directly with oxygen, but also with hydrogen, forming a compound called acetylene, C_2H_2. Carbon forms with oxygen, hydrogen, and nitrogen a series of more or less complicated compounds very much more extended than the series formed with these bodies by any other element,

so that these compounds are considered as a separate branch of the science under the name of Organic Chemistry.

The Diamond was first found to consist of pure carbon by Lavoisier, in 1775-6, by burning it in oxygen, and collecting the carbon dioxide formed, it occurs crystallized in certain sedimentary rocks and gravel in India, Borneo, and Brazil. Diamond occurs crystallized in forms derived by a symmetrical geometric operation from a regular octahedron. The specific gravity of diamond varies from 3.3 to 3.5, it is the hardest of all known bodies, and when cut possesses a brilliant lustre, and a high refractive power. In addition to its employment as a gem, the diamond is used for cutting and writing upon glass. We are altogether unacquainted with the mode in which the diamond has been formed, it cannot, however, have been produced at a high temperature, because, when heated strongly in a medium incapable of acting chemically upon it, the diamond swells up, and is converted into a black mass resembling coke.

Graphite, or Plumbago, crystallizes in six-sided plates which have no relation to the form in which the diamond crystallizes. Graphite occurs in the oldest sedimentary formations, and in granitic or primitive rocks, it is found in large quantities in Siberia and Ceylon. It has a black metallic appearance, whence the familiar name black lead, and leaves a mark when drawn upon paper. The specific gravity of graphite is 2.15 to 2.35. Coarse impure graphite may be purified by heating the powder with sulphuric acid and potassium chlorate, a compound is thus obtained which, on being heated strongly, decomposes, leaving pure

graphite in a bulky and finely-divided powder. This powder when strongly compressed forms a coherent mass, from which pencils and other articles can be made. Graphite is used for polishing surfaces of iron-work, and also for giving a protecting varnish to grains of gunpowder. Graphite is produced in the manufacture of iron, it occasionally separates from the molten pig-iron in the form of scales.

Charcoal is the third allotropic modification of carbon. It is obtained in a more or less pure state whenever animal or vegetable matter is heated to redness in a vessel nearly closed, the volatile matters, compounds of carbon, hydrogen, and oxygen, are thus driven off, and the residue of the carbon, together with the ash or mineral portion of the organism, remains behind.

The purest form of charcoal-carbon is found in lamp-black, it also occurs as wood charcoal, coal, coke, and animal charcoal. This form of carbon does not crystallize, and is hence termed amorphous carbon. It is much lighter than either of the other two forms, the specific gravity of powdered coke varying from 1.6 to 2.0. Charcoal appears at first sight to be lighter than water, as a piece of it floats on the surface of this liquid, this is, however, due to the porous nature of the charcoal, for if it be finely powdered it sinks to the bottom of the water. This porous nature of charcoal enables it to exert a remarkable absorptive power, of which much use is made in the arts. Charcoal is thus able to absorb about ninety times its own volume of ammonia gas, and about nine volumes of oxygen. In the process of sugar-refining, use is made of the property of charcoal to absorb the coloring mat-

ters present in the raw sugar. The kind of charcoal best suited to this purpose is that obtained by heating bones in a closed vessel. Charcoal is also used as a disinfectant in hospitals and dissecting rooms. It appears that the putrefactive gases when absorbed by the charcoal undergo a gradual oxidation from contact with the oxygen of the air taken up by the charcoal, and are thus rendered harmless.

Coal is a form of carbon less pure than wood charcoal. It consists of the remains of a vegetable world which once flourished on the earth's surface. The original woody fibre has undergone a remarkable transformation in passing into coal, having been subjected to a process similar, in a chemical point of view, to that by which wood is transformed into charcoal. It has not, however, lost the whole of its hydrogen and oxygen, and it has at the same time become bitumenized, so that for the most part all the vegetable structure has disappeared. There are many different kinds of coal, containing more or less of the oxygen and hydrogen of the original wood. Cannel coal and bog coal contain the most hydrogen, and anthracite coal the least.

Carbon and Hydrogen Compounds.

Methane or Carburetted Hydrogen, CH_4, Acetylene, C_2H_2 and Ethylene or Olefiant Gas, C_2H_4.

These compounds are very numerous, they are known in the gaseous, liquid, and solid forms. A still larger number of substances exist containing carbon, hydrogen, and oxygen, with sometimes nitrogen, these are termed organic compounds, and they are more numerous than all the compounds of the other elements

put together. Many of these are found to be formed from the bodies of plants and animals.

Methane or Marsh Gas: Formula—CH_4. Molecular Weight—15.9.

This is a colorless, tasteless, inodorous gas, which has not been condensed to a liquid. It is found in coal mines, and known under the name of firedamp, it also occurs in stagnant pools, being produced by the decomposition of dead leaves, whence the name marsh gas. It is one of the constituents of coal gas, etc., and is evolved in many volcanic districts. Marsh gas may also be artificially prepared by heating sodium acetate with caustic soda.

Sodium acetate and caustic soda give sodium carbonate and marsh gas.

Marsh gas burns with a bluish-yellow, non-luminous flame, forming carbon dioxide and water; with a limited supply of air it yields several products, amongst which is acetylene. If mixed with ten times its volume of air, or twice its volume of oxygen, it ignites with a sudden and violent explosion on the application of a light, and hence the great damage produced by the escape of this gas in coal mines. The composition of marsh gas is ascertained by exploding it with oxygen in the eudiometer. 1 volume of this gas and 3 volumes of oxygen yield 2 volumes after passage of the spark. On absorbing by potash the carbon dioxide produced, 1 volume of oxygen is found to remain. Hence of the 2 volumes of oxygen needed to burn the 1 volume of marsh gas, 1 has gone to unite with the carbon, and 1 to form water with the hydrogen. It is thus seen that 2 volumes of marsh gas contain 4 volumes of hydrogen

weighing 4, as water contains 2 volumes of hydrogen and 1 of oxygen, and as much carbon as is contained in 2 volumes of carbon dioxide, viz., 12 parts by weight, and hence the formula CH_4 is given to this gas.

Acetylene: Formula—C_2H_2.

This gas is formed by the direct union of carbon and hydrogen at a very high temperature. For this purpose the carbon terminals of a powerful galvanic battery are brought together in an atmosphere of hydrogen. At the high temperature thus evolved, a direct union of carbon and hydrogen takes place, and acetylene is formed. Acetylene is a colorless gas, which burns with a bright luminous flame, and possesses a disagreeable and very peculiar odor, it is produced in all cases of incomplete combustion, and its smell may be noticed when a candle burns with a smoky flame. Acetylene combines with certain metals, such as copper and silver, and the compounds thus formed are distinguished by the ease with which they undergo explosive decomposition. This gas likewise unites directly with hydrogen, forming the next substance, ethylene, $C_2H_2+H_2=C_2H_4$.

Acetylene is now manufactured on a large scale for use as an illuminant in places where coal-gas is not available, such as small villages and country houses or railway stations. The method invariably adopted for its production is the action of water on calcium carbide. The carbide was, however, first obtained in quantity in 1892 by a Canadian engineer named Willson, in an attempt to manufacture metallic calcium by heating lime with carbonaceous matter in the elec-

tric furnace, the resulting product being, however, calcium carbide.

The manufacture of acetylene from carbide was commenced about 1895, and since that date some hundreds of different forms of apparatus have been patented for carrying out the decomposition of the carbide with water. These may all be divided into two general classes: Those in which the water is added to the carbide, and those in which the carbide is added to the water so that the latter is always in excess. The objection to generators of the first class is the higher temperature attained in the reaction, which results in the partial conversion of acetylene into tarry matters. Some generators of each class are constructed so as to render them "automatic," the generation of acetylene only takes place as fast as it is consumed, whilst in others the whole charge is decomposed at once, and the resulting gas stored in a holder for use as required.

Ethylene or Olefiant Gas: Formula—C_2H_4.

This gas is obtained on the destructive distillation of coal, and is an important constituent of coal gas. It is obtained in the pure state by heating 1 part of alcohol (spirits of wine), C_2H_6O, with 5 or 6 parts by weight of strong sulphuric acid, as in formation of carbon monoxide from formic acid the elements of water are separated by the sulphuric acid, and C_2H_4 is evolved as a gas. This gas is colorless, but possesses a sweetish taste, by exposure to a high pressure at a temperature of —110° it has been condensed to a colorless liquid. On bringing it in contact with a light in the air, it burns with a luminous smoky flame, forming carbon dioxide and water. When mixed

with three times its bulk of oxygen and fired, it detonates very powerfully. 1 volume of olefiant gas requires 3 volumes of oxygen to burn it completely, and yields 2 volumes of carbon dioxide, so that 1 volume of oxygen is needed to combine with the hydrogen. Hence this gas contains twice as much carbon as marsh gas, with the same quantity of hydrogen, we must therefore write its formula C_2H_4.

Olefiant gas combines directly with its own volume of chlorine gas, forming an oily liquid, $C_2H_4Cl_2$; and owing to this property it has received the above name.

Carbon, Hydrogen and Oxygen Compounds.

Ether or Diethyl Ether: Formula—$C_4H_{10}O$.

Ether is prepared on a large scale by heating a mixture of alcohol and sulphuric acid to 140°, when ether and water are given off. The decompositions which take place are as follows: In the first place, alcohol and sulphuric acid form hydrogen-ethyl-sulphate (sulphovinic acid) and water, by an exchange of hydrogen and ethyl.

This hydrogen-ethyl-sulphate next comes in contact with a second molecule of alcohol, another exchange of hydrogen for ethyl occurs, and ether and sulphuric acid are formed.

The water formed by the first decomposition, and the ether produced by the second, are given off as vapor, whilst the sulphuric acid remains behind, ready again to go through the same series of changes on meeting with two other molecules of alcohol. This process is called the continuous etherification process, as a current of alcohol may be passed continuously

through the sulphuric acid heated to 140°, and a regular supply of ether and water thus obtained.

Ether is a colorless, very mobile liquid, possessing a strong and peculiar smell. It is lighter than water, and is not miscible with that liquid. Ether boils at 34°, and its vapor is 37 times heavier than hydrogen, and can be poured from vessel to vessel like carbonic acid gas. It burns with a luminous flame, and explodes when mixed with air. From its low boiling point great care must be taken to avoid explosions when working with this substance, owing to the vapor becoming mixed with air. Ether is easily attacked by oxidizing agents, yielding the same products as alcohol, it is also acted upon by chlorine.

Carbon and Nitrogen Compounds.

Cyanogen Gas: Formula—C_2N_2.

This gas is formed when the cyanides of mercury, silver, or gold are heated. For this purpose it is usual to employ mercuric cyanide, which decomposes thus

$$Hg(CN)_2 = Hg + C_2N_2.$$

The mercury salt, placed in a tube of hard glass fitted with a cork and gas-delivery tube, or in a small hard glass retort, is heated to dull redness, and the gas collected over mercury.

The heat of formation of mercuric cyanide is large, and a very high temperature is therefore required to bring about this decomposition. If mercuric chloride be mixed with the cyanide, the cyanogen comes off at a lower temperature, as the mercury then combines with the mercuric chloride forming mercurous chlor-

ide, and the whole reaction, represented by the equation

$$Hg(CN)_2+HgCl_2=(CN)_2+2HgCl.$$

It is also formed by gently igniting an intimate mixture of two parts of well-dried potassium ferrocyanide, $Fe(CN)_6K_4$, with three parts of mercuric chloride. The gas thus prepared contains free nitrogen, and the employment of potassium cyanide in place of the ferrocyanide is recommended. There is no doubt that in this reaction mercuric cyanide is first obtained, and this is then decomposed as already described.

Cyanogen is most readily obtained by gradually adding a concentrated solution of potassium cyanide to a solution of two parts of crystallized copper sulphate in four parts of water, and then warming, cyanogen is evolved and cuprous cyanide simultaneously separates out from the solution. If the cuprous cyanide be washed by decantation, and then treated with ferric chloride, the remainder of the cyanogen is evolved.

Cyanogen gas is found in small quantities in the gases proceeding from the blast furnaces, and it is likewise formed when a mixture of ammonia and coal-gas is burnt in a Bunsen burner. The following equation probably explains the reaction which takes place in the latter case.

$$4CO+4NH_3+O_2=2C_2N_2+6H_2O.$$

Cyanogen is a colorless gas possessing a peculiar pungent odor resembling that of peach kernels. It is poisonous, and burns, when ignited in air, with a characteristic purple-mantled flame, with formation of carbon dioxide and nitrogen.

Carbon, Hydrogen and Chlorine Compounds.

Chloroform: Formula—$CHCl_3$.

Chloroform is formed when chlorine acts on marsh gas, but it is prepared by acting upon methyl or ethyl alcohol with bleaching powder. It is a mobile, heavy liquid, possessing a powerful and agreeable smell. Chloroform is much used in medicine, producing, when it is inhaled, a temporary but perfect insensibility to pain, and is therefore much valued in surgical operations.

Carbon, Hydrogen and Oxygen Compounds.

Methyl or Wood Alcohol: Formula—$CH_3.OH$.

Methyl alcohol is produced by the dry distillation of wood, forming about one per cent of the aqueous distillate. It is likewise met with in the oil of wintergreen. Methyl alcohol can likewise be synthetically built up from its constituent elements by means of several complicated reactions.

Pure methyl alcohol is obtained from crude wood alcohol, in which it is contained mixed with a variety of other organic compounds, by forming a crystalline methyl oxalate. This, on treatment with water, is decomposed, and yields the alcohol in the pure state. Methyl alcohol is a colorless, mobile liquid, possessing a pure spirituous smell. It burns with a non-luminous flame, and is soluble in and miscible with water. Potassium dissolves in methyl alcohol with evolution of hydrogen and formation of potassium methylate. Methyl alcohol when acted on by oxidizing agents yields methyl aldehyde and formic acid. By

the action of bleaching-powder on methyl alcohol, chloroform is obtained, acted upon by hydrochloric acid, the alcohol yields methyl chloride.

Ethyl or Absolute Alcohol: Formula—$C_2H_5.OH$.

Ethyl Alcohol is obtained in the vinous fermentation of sugar, a decomposition effected in aqueous sugar solutions in presence of yeast, in which alcohol and carbonic acid are chiefly formed.

Alcohol and alcoholic liquids are prepared in large quantities by the fermentation of sugar derived from various sources. The fermented liquid is distilled, and the dilute aqueous spirit thus separated from non-volatile impurities. It is obtained in a more concentrated form by repeated rectifications, as it boils at a lower temperature than water. Alcohol cannot, however, be completely separated from water by simple distillation, the strongest spirit which can thus be prepared containing 10 per cent of water. To withdraw all the water, the spirit must be distilled with some substance capable of combining with water, such as potassium carbonate or quicklime. The pure liquid thus obtained is termed **absolute alcohol.** It is a colorless, mobile liquid, possessing a pleasant, spirituous smell and burning taste and it boils at 78°.4 when the barometer stands at 760 mms. It has not been solidified, becoming only viscid at a temperature of —100°. Alcohol is very inflammable, burning with a slightly luminous blue flame. It absorbs moisture with great avidity, and mixes with water in all proportions, the mixture evolving heat and undergoing contraction in volume.

Carbon and Oxygen Compounds.

Carbon Dioxide, CO_2 and Carbon Monoxide, CO.

Carbon Dioxide or Carbonic Acid Gas: Formula—CO_2. Molecular Weight—43.7.

The carbonic acid contained in the air is derived from a variety of sources, it is formed by the respiration of man and animals, as well as in the act of combustion of organized material, and in its decay and decomposition. The amount of atmospheric carbonic acid varies between certain narrow limits, but on an average reaches 3 volumes in 10,000 volumes of air. In the presence of the sunlight, plants have the power, through their leaves, of decomposing this carbonic acid, taking up the carbon to form their own tissue, and eliminating the oxygen gas, hence the amount of carbonic acid in the air does not increase beyond the limits named.

Carbon dioxide is an acid-forming oxide giving rise to a series of salts termed the carbonates, many of which occur in nature as minerals. Amongst these is especially to be mentioned calcium carbonate, $CaCO_3$, which occurs in two distinct crystalline forms as calcspar and arragonite, while it is found in a massive crystalline form in marble and limestone, calcium carbonate also forms the chief constituent of the shells of mollusca and foraminifera, the remains of which constitute the chalk formation as well as the greater part of all the limestones. In order to prepare carbon dioxide, a carbonate, such as marble or chalk, is brought into a gas-evolution flask, and dilute hydrochloric acid

poured upon it, when the gas is rapidly evolved with effervescence, thus:

$$CaCO_3+2HCl=CO_2+CaCl_2+H_2O.$$

The gas thus obtained invariably carries over small quantities of hydrochloric acid vapor with it, from which it may be freed by passing through a solution of sodium bicarbonate. In order to obtain a constant stream of carbon dioxide the same apparatus may be employed which is made use of for the preparation of sulphuretted hydrogen gas.

Another method of obtaining a constant current is to pour concentrated sulphuric acid over chalk, and add a very small quantity of water, thus:

$$CaCO_3+H_2SO_4=CO_2+CaSO_4+H_2O.$$

The dilute acid cannot be employed for this purpose, since the calcium sulphate formed, which is soluble in the concentrated acid, does not dissolve in the dilute acid, and thus prevents its further action on the chalk.

The gas obtained from chalk possesses a peculiar smell, which is due to the presence of small quantities of volatile organic matter always contained in the chalk. In order to prepare a very pure gas, sodium carbonate may be decomposed with pure dilute sulphuric acid, thus:

$$Na_2CO_3+H_2SO_4=CO_2+Na_2SO_4+H_2O.$$

Carbon dioxide may be obtained on a large scale for the preparation of sodium bicarbonate, white lead, and other commercial products by burning limestone, or by the combustion of charcoal or coke, but the gas then always contains large volumes of nitrogen derived from the air used for combustion. In order to obtain the dioxide in comparatively pure condition

from such gases on the large scale, the mixture may be passed into a concentrated solution of potassium carbonate with which carbon dioxide combines forming potassium hydrogen carbonate, $KHCO_3$. The solution of the latter evolves carbon dioxide on warming, leaving a solution of the normal carbonate, which is used again for absorbing fresh quantities of the dioxide.

Carbon Monoxide or Carbonic Oxide Gas: Formula—CO. Molecular Weight—27.8.

When carbon burns with a limited supply of oxygen, carbonic oxide is formed. The production of this gas in an ordinary red-hot coal fire is often observed. Oxygen of the air, which enters at the bottom of the grate, combines with the carbon of the coal, forming carbon dioxide, this substance then passing upwards over the red-hot coals, parts with half its oxygen to the red-hot carbon, thus

$$CO_2+C=2CO.$$

This carbon monoxide on coming out at the top of the fire meets with atmospheric oxygen, with which it at once combines, burning with a lambent blue flame, and re-forming carbon dioxide. Carbon monoxide can be prepared in various ways. It is formed when zinc oxide, ferric oxide, manganese dioxide, and many other oxides are heated with charcoal, it is also formed when chalk (calcium carbonate), magnesite (magnesium carbonate), and other carbonates are heated with metallic zinc or iron filings, the decomposition which takes place in these cases is represented by the following equations,

$$ZnO+C=Zn+CO.$$

$$CaCO_3+Zn=CaO+ZnO+CO.$$

Carbon monoxide is also formed when carbon is

burnt in a limited quantity of oxygen, or by the action of carbon dioxide on heated charcoal.

$$CO_2+C=2CO.$$

The latter reaction does not take place below 600°, and then only in presence of moisture. As the temperature is increased above 600°, moist carbon dioxide is partially converted into the monoxide, an equilibrium between the two gases being obtained at any particular temperature.

When oxalic acid or an oxalate is heated with concentrated sulphuric acid, a mixture of equal volumes of carbon monoxide and carbon dioxide is evolved, thus

$$C_2O_4H_2=H_2O+CO+CO_2.$$

The carbon dioxide may readily be separated from the carbon monoxide either by passing the mixed gases through a solution of caustic soda or by collecting the mixture over water rendered alkaline by this substance.

Carbon and Sulphur Compounds.

Carbon Bisulphide: Formula—CS_2. Molecular Weight —75.57.

Carbon bisulphide is prepared on a large scale by passing the vapor of sulphur over red-hot charcoal. For this purpose a large upright cast iron cylinder is employed, 10 or 12 feet long and one to two feet in diameter as in Fig. 8. This cylinder is placed above a furnace and surrounded by brickwork, and at the same time it is provided with a lid to admit of the whole being filled with charcoal. A second opening **a**, furnished with a hopper, exists at the bottom of the cylinder, and this serves to bring the sulphur into the

apparatus. The sulphur evaporates, and in the state of vapor combines with the red-hot carbon, impure carbon bisulphide distilling over by the tube **c**, and collecting in the vessel **d** under water. The tubes **e** serve as condensers, to separate the vapor of carbon bisulphide from sulphuretted hydrogen formed during the reaction, owing to the presence of hydrogen in the charcoal, the sulphuretted hydrogen being absorbed by passing over the layers of slaked lime con-

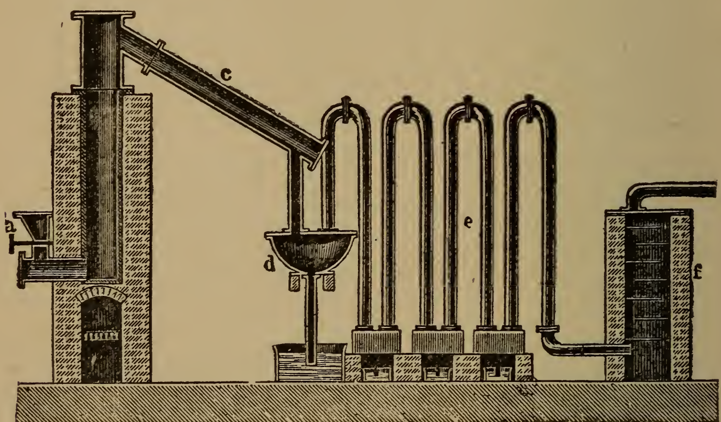

Fig. 8.

tained in the purifier f. In actual work the yield of bisulphide is about 20 per cent. below the theoretical amount. The crude substance invariably contains sulphur in solution, from which, however, it may be separated by distillation, but other sulphur compounds, which impart to the crude material a most offensive odor, are also contained in the distillate. In order to remove these impurities, different processes are in use. The substance was formerly purified by

frequent re-distillation over oil or fat, by which means the disagreeably smelling compounds were held back. Another means of purification employed is that of shaking the liquid with mercury, and allowing it to remain for a long time in contact with corrosive sublimate in the cold, and then distilling it off white wax.

Pure carbon bisulphide is a colorless, mobile, strongly refracting liquid, possessing a sweetish smell not unlike that of ether or chloroform. When carbon combines with sulphur to form carbon bisulphide, heat is absorbed which is given out again when it is decomposed. Like many substances of this class it is decomposed into its elements when the vapor is subjected to the shock caused by the explosion of fulminate of mercury, but at atmospheric pressure the decomposition is only local and is not transmitted through the whole mass of the vapor.

Carbon bisulphide is very inflammable and burns in the air, forming carbon dioxide and sulphur dioxide. The combination with oxygen takes place slowly below the temperature of ignition, a slight phosphorescence being observed, and as usual in such cases the exact temperature of ignition varies, but it occurs immediately at temperatures above 260°. The vapor of carbon bisulphide when mixed with oxygen can be exploded by a spark even when the mixture is thoroughly dried, differing in this respect from carbon monoxide. With excess of oxygen a mixture of carbon dioxide, sulphur dioxide, and trioxide, and sometimes free sulphur is produced, carbon monoxide, carbonyl sulphide, and unaltered carbon bisulphide being also present in the explosion wave. With an insufficient quantity of oxygen, the products are carbon

dioxide and monoxide, sulphur dioxide, carbonyl sulphide and unaltered carbon bisulphide, but in no case is free carbon formed, and it is therefore improbable that the vapor is dissociated into its elements before combining with oxygen.

CHLORINE.

Symbol—Cl. Atomic Weight—35.2. Density—35.2.

Chlorine does not occur free in nature, but can easily be prepared from its compounds. It is found combined with metals forming chlorides, of these sodium chloride, sea- or rock-salt, is the most common. To obtain chlorine from this, it must be heated with sulphuric acid and manganese dioxide.

Sodium chloride, sulphuric acid, and manganese dioxide give chlorine, sodium sulphate, manganese sulphate, and water.

If one part by weight of salt to one part of manganese dioxide be mixed with two parts of sulphuric acid and two of water, and the mixture brought into a large flask, the chlorine gas is given off regularly upon the application of a very slight heat. In order to obtain the gas pure, it may be passed through the water contained in a wash-bottle before it is collected for use. Chlorine is a green-yellow gas, possessing a most disagreeable and peculiar smell, which, when the gas is present in small traces only, resembles that of seaweed, but when present in large quantities acts as a violent irritant, producing inflammation of the mucous membrane, and even causing death when inhaled. Chlorine gas when submitted to a pressure of five at-

mospheres at the ordinary temperature is condensed to a heavy yellow liquid, but it has not yet been solidified. This gas cannot be collected over water or mercury, as it is soluble in the former, 1 volume of water dissolving 2.37 volumes of chlorine at 15°, and it combines directly with the latter, forming mercuric chloride. It can, however, be easily collected by displacement, as it is nearly 2.5 times as heavy as air. If metals in a freely divided state are brought into chlorine gas, they take fire spontaneously, forming metallic chlorides. Powdered arsenic, antimony, or thin copper leaf, burn when thrown into the gas.

The most remarkable property of chlorine is its power of combining with hydrogen to form hydrochloric acid. When these two gases are brought together in equal volumes, they combine with explosion on bringing a flame into contact with them, or on exposing the mixture to sunlight. Chlorine is even able to decompose water in the sunlight, combining with the hydrogen and liberating the oxygen. Several experiments illustrative of this property of chlorine may be mentioned. If a burning candle be plunged into this gas, the taper continues to burn, but with a very smoky flame, the hydrogen alone of the wax entering into combination with the chlorine, whilst the carbon is given off as smoke or soot. The same effect is produced when a paper moistened with turpentine, a compound of carbon and hydrogen, is held in a jar of chlorine gas, the hydrogen of the turpentine at once combines with the chlorine, forming hydrochloric acid, and the carbon is liberated, so much heat is given off by this action that the paper frequently takes fire.

The well-known bleaching action of chlorine also de-

pends upon its power of combining with the hydrogen of water and liberating the oxygen. Dry chlorine gas does not bleach, we may enclose a piece of cotton cloth or paper colored by a vegetable substance, as madder or indigo, in a bottle of dry chlorine, and no change of color takes place, even after the lapse of many weeks. If a few drops of water are added, the coloring matter is immediately destroyed, and the cotton or paper is bleached. Here the chlorine combines with the hydrogen of the water, and the oxygen at the moment of its liberation combines with the vegetable coloring matters, forming compounds destitute of color. Ordinary free oxygen has not this power, not at least to any great extent, and it is a frequent observation that bodies in this nascent state have more active properties than the same bodies when in the free state. This difference depends upon the fact that the molecules, or smallest particles of an element which can exist in the free state, do not consist of the individual atoms, but of a group of atoms. The molecule of a compound body contains two or more dissimilar atoms, whilst that of an element contains similar atoms. The moment an element is liberated from a compound, the single atoms unite together to form a molecule, and the elementary body makes its appearance in the free state. If substances are present on which the element can act chemically, they are decomposed by the chemical attractions of the liberated atom, which are more active in that state than when united to form a molecule.

Chlorine is unable to bleach mineral colors. The difference between printers' ink, colored by lampblack or carbon, and writing ink, a vegetable black, is well

illustrated by placing a sheet of paper having characters written and printed upon it in a solution of chlorine in water. Chlorine gas is largely used for bleaching purposes in the cotton, linen, and paper manufactures. It is sometimes used in the form of a gas, but more usually in combination with calcium and oxygen, forming the article called chloride of lime, a mixture of calcium chloride $CaCl_2$, and calcium hypochlorite, $CaCl_2O_2$, or bleaching powder. Chlorine is also largely employed as a disinfectant and deodorant, its action on organic putrefactive substances being similar to that upon organic coloring matters.

Chlorine is usually prepared when a high degree of purity is not required, by the action of black oxide of manganese (manganese dioxide) on strong hydrochloric acid, thus

$$MnO_2+4HCl=Cl_2+MnCl_2+2H_2O.$$

When these two substances are brought together, a dark greenish-brown solution is first obtained, and this on heating evolves chlorine gas, whilst manganous chloride $MnCl_2$ is formed. There is little doubt that the dark-colored solution contains a higher unstable manganese chloride, probably $MnCl_3$ or $MnCl_4$, which has not yet been isolated, and which on heating decomposes into chlorine and manganous chloride.

In preparing the gas by this method the oxide of manganese should be in the form of small lumps free from powder, and the hydrochloric acid poured on so as to about cover the solid, on gently warming, the gas is copiously evolved.

It is often more convenient for laboratory uses to liberate the hydrochloric acid in the same vessel in

which it is acted upon by the manganese dioxide, and for this purpose to place a mixture of one part of the latter substance and one part of common salt in a large flask as in Fig. 9, containing a cold mixture of two parts of strong sulphuric acid and two of water, on very slightly warming the mixture, a regular evolution of gas takes place.

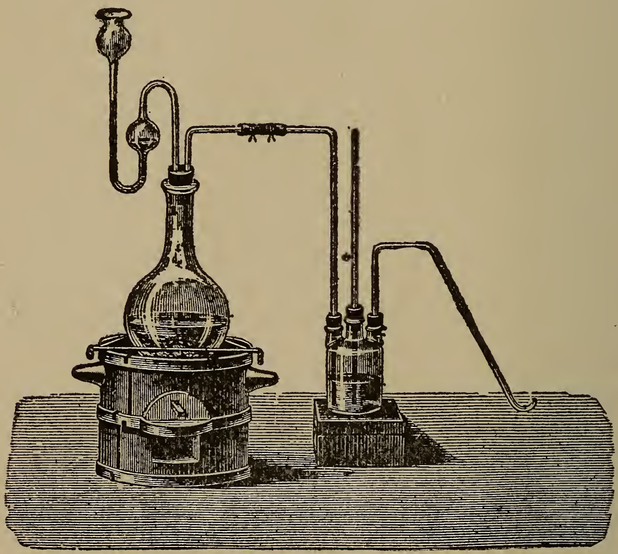

Fig. 9.

Chlorine is not inflammable, and does not directly combine with oxygen, it unites, however, with great energy with hydrogen, forming hydrochloric acid, HCl, and to this property it owes its peculiar and valuable bleaching power. It also combines with many metals, giving rise to a class of compounds termed the metallic chlorides.

In each case of combination with chlorine a definite

quantity of heat is given out, whilst sometimes light is also emitted, so that the essential phenomena of combustion are observed. Thus if we plunge a jet from which a flame of hydrogen burns into a cylinder of chlorine gas as in Fig. 10, the hydrogen continues to burn, but instead of water being produced, hydrochloric acid is formed by the combustion. In like

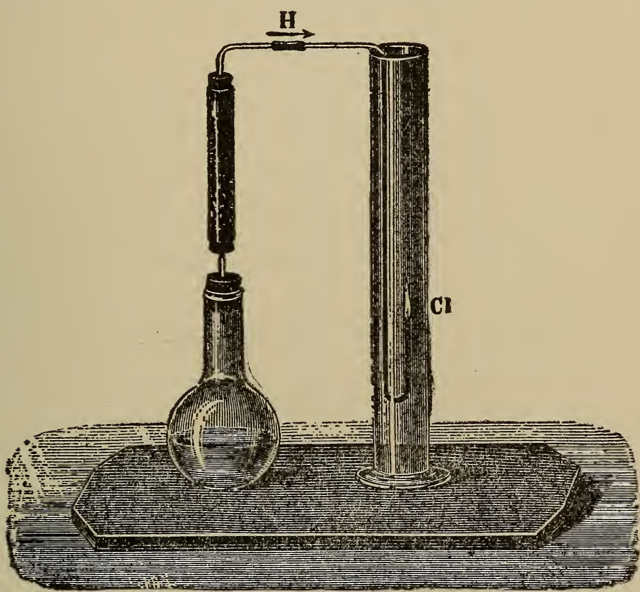

Fig. 10.

manner, if we bring a light to the mouth (held downwards) of a cylinder of hydrogen as in Fig. 11, and then bring this over a jet from which chlorine gas is issuing, a flame of chlorine burning in hydrogen will be seen.

If two equal sized cylinders, filled, one with chlorine and the other with hydrogen, are brought mouth to

mouth, the two glass plates closing them withdrawn, and the gases allowed to mix, and if then a flame is brought near the mouths of the cylinders, the mixed gases combine with a peculiar noise, and dense fumes

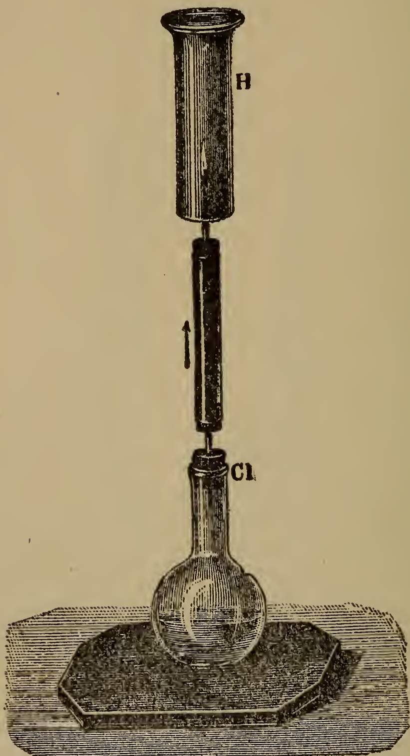

Fig. 11.

of hydrochloric acid gas are seen. This experiment must, however, be made in a room partially darkened, or performed by gas or candle-light, as the two gases combine with explosion in sunlight or strong daylight.

CHROMIUM.

Symbol—Cr. **Atomic Weight—51.7.**

Chromium is a substance whose compounds do not occur very widely distributed, or in large quantities, but they are much employed in the arts as pigments, many of them possessing a fine bright color. The chief ore of this metal is Chrome Ironstone, a compound isomorphous with Magnetic Oxide of Iron, found in America, Sweden, and the Shetlands, a compound lead chromate is also found in some quantity. Pure chromium appears to be the most infusible of all the metals, as it cannot be melted at a temperature sufficient to fuse and volatilize platinum. It has been obtained by another process in the form of bright crystals.

Chromium and Chlorine Compounds.

Chromic Chloride: Formula—Cr_2Cl_6.

Chromic Chloride, the anhydrous chloride, is obtained as a sublimate, in beautiful violet crystals, by passing a current of chlorine gas over a red-hot mixture of chromium sesquioxide and charcoal. These crystals do not dissolve easily in water, but are readily soluble if a trace of chromium dichloride is present. The most ready way of preparing a solution of chromic chloride is to boil a solution of chromic acid or a chromate with hydrochloric acid and alcohol, the red or yellow solution after a few minutes being changed to a deep greenish-blue color. A solution of chromic sulphate may be obtained in the same way, by substituting sulphuric acid for hydrochloric acid. Chromium sulphate forms a series of alums with potassium

and ammonium sulphates, which have a deep purple tint, and are isomorphous with common alum.

Chromium and Oxygen Compounds.

Chromium Monoxide, CrO, Chromium Sesquioxide, Cr_2O_3 and Chromium Trioxide, CrO_3.

Chromium Monoxide: Formula—CrO.

Chromium Monoxide is only known in the hydrated state, as both it and its compounds absorb oxygen with great avidity. The hydrate is prepared as a brown precipitate by adding potash to the solution of chromium dichloride.

Chromium Sesquioxide: Formula—Cr_2O_3.

Chromium Sesquioxide, or Chromic Oxide is a dark green, perfectly stable powder, obtained by igniting the hydroxide formed by precipitating any soluble chromic salt with ammonia. It is employed as a green color for painting on porcelain, and produces the green of the emerald. A splendid green color is also obtained by heating potassium bichromate with boron-trioxide, on dissolving in water a grass-green hydroxide remains behind, which is termed Guignet's green.

Chromium Trioxide: Formula—CrO_3.

Chromium Trioxide is obtained in the form of long ruby-red needle-shaped crystals by adding an excess of strong sulphuric acid to a concentrated solution of the bichromate. The crystals are very soluble in water, forming an acid solution of chromic acid. The excess of sulphuric acid may be removed by washing with concentrated nitric acid, and the crystals then

dried in a current of air in a glass tube. The crystals of chromium trioxide are very easily reduced to sesquioxide in presence of organic matter. So energetic is this evolution of oxygen, that ignition occurs when alcohol is dropped on the dry crystals.

If a solution of chromium trioxide or of potassium bichromate is heated with hydrochloric acid, chromic chloride is formed and chlorine liberated, whereas, if chromium trioxide is heated with sulphuric acid, a chromic sulphate is formed and oxygen gas is given off.

Chromic Acid and the Chromates.

If any chromic compound be fused with potassium carbonate, it becomes oxidized, and a soluble yellow chromate is formed, K_2CrO_4. This is the mode in which the chromium compounds are prepared from chrome-iron ore. This yellow chromate is isomorphous with potassium, sulphate and manganate. When sulphuric acid is added to a solution of this yellow salt in sufficient quantity to combine with half the base, large red crystals of the bichromate, $K_2Cr_2O_7$, separate out. This salt is largely used for the preparation of the chrome pigments. If to the solution of the bichromate a solution of chromium trioxide be added, a third salt, termed ter-chromate, $K_2Cr_3O_{10}$, crystallizes out.

COBALT.

Symbol—Co. **Atomic Weight—59.0.**

Cobalt is a reddish-white, very tenacious metal, which is as infusible as iron, and, like the latter metal, is strongly magnetic. It is not found native, but occurs in combination with arsenic and sulphur, as two

distinct minerals. The metal dissolves slowly in sulphuric and hydrochloric acids with evolution of hydrogen. The cobalt compounds are distinguished for the brilliancy of their color, they are employed as pigments, and they impart a magnificent blue tint to glass.

The nitrate and sulphate of cobalt are also soluble salts, the latter is isomorphous with magnesium sulphate. Cobalt sulphide is a black powder, insoluble in dilute acids. Cobalt compounds can be easily recognized by the deep blue tint which very minute traces impart to glass, or to a borax bead, made by fusing borax into a colorless mass on the loop of a platinum wire.

Cobalt and Chlorine Compounds.

Cobalt Chloride: Formula—$CoCl_2$.

Cobalt chloride is a soluble salt obtained by acting on the oxide or on the metallic ore with hydrochloric acid, the solution yields on evaporation pink crystals of the hydrated chloride, or, if further heated, blue crystals of the anhydrous salt.

Cobalt and Oxygen Compounds.

There are three oxides of cobalt, the monoxide, CoO, the sesquioxide, Co_2O_3 and an oxide Co_3O_4, the former, on solution in acids, forms the series of cobalt salts, which are pink when hydrated, and blue when anhydrous, while the sesquioxide does not form any salts. Cobalt monoxide, CoO, is obtained as a brown powder by carefully heating the rose-colored hydrate, precipitated by potash in solutions of cobalt, and Cobalt

sesquioxide, Co_2O_3, is prepared by adding a solution of bleaching powder to a soluble protosalt.

COPPER.

Symbol—Cu. **Atomic Weight—63.13.**

Copper is an important metal, largely used in the arts, and has been known from early times, as it occurs native in the metallic state, and is moreover easily reduced from its ores. Metallic copper is found in considerable quantity in America and other localities, but the chief sources of copper are the following ores. A compound of copper, sulphur, and iron, known as copper pyrites, the cuprous sulphide, the carbonate or malachite, and the red or cuprous oxide. Pure metallic copper can be obtained by reducing the oxide in a current of hydrogen gas, or by the electrolytic decomposition of a salt of copper. The process for obtaining copper on a large scale from the carbonate or oxide is a very simple one, merely reducing these ores together with carbon and some silica in a furnace. The reduction of the metal is more difficult when the commoner ore, copper pyrites, is employed. In this case the ore is repeatedly roasted, in order partially to convert the cuprous sulphide into oxide, and the roasted ore melted in a furnace with the addition of sand or silicious slag, in this operation the cuprous oxide becomes converted into the corresponding sulphide, whilst the iron oxidizes and unites with the silica to form a light and fusible slag. The impure cuprous sulphide fuses and sinks to the lower portion of the furnace, forming the mat or coarse metal, and by repeating this operation a pure cuprous sulphide or fine metal is obtained. In order to prepare the metal-

lic copper free from sulphur, this fine metal is roasted, and afterwards fused in contact with the air. During the first part of the operation a portion of the sulphur is burnt off, cupric oxide being formed, and in the later stages of the process this oxide acts upon the remaining quantity of sulphide, forming sulphur dioxide and metallic copper. In order to get rid of the last traces of oxide, the molten copper is poled or stirred up with a piece of green wood.

Metallic copper possesses a peculiar deep red color, which is best seen when a ray of light is several times reflected from a bright surface of the metal, it is very malleable and ductile, and possesses great tenacity, it melts at a red heat, and is lightly volatile at a white heat, communicating a green tint to a flame of hydrogen gas, and it is one of the best conductors of heat and electricity. Copper does not oxidize either in pure dry or moist air at ordinary temperatures, but if heated to redness in the air, it rapidly oxidizes to scales of copper oxide. Steam is not decomposed by metallic copper at a red heat. Finely divided copper dissolves in hydrochloric acid with evolution of hydrogen, when heated with strong sulphuric acid, sulphur dioxide is evolved, and copper sulphate formed, and when acted upon with nitric acid, copper nitrate is produced, and nitric oxide liberated.

Many of the copper alloys are of importance. Brass is an alloy containing about two-thirds of copper and one-third of zinc, it is harder than copper, and can be more easily worked. The addition of one to two per cent of lead improves the quality of brass for most purposes. Bronze, gun-metal, bell-metal, and speculum-metal are other alloys of copper and tin in

varying quantities. They are all remarkable for the property of being hard and brittle when slowly cooled, but of becoming soft and malleable if they are cooled suddenly when red-hot by dipping into cold water.

Copper and Chlorine Compounds.

Chloride of Copper: Formula—$CuCl_2$.

Copper chloride is formed when copper is brought into chlorine gas, or when copper oxide is dissolved in hydrochloric acid; it forms green needle-shaped crystals, soluble in water and alcohol. The alcoholic solution burns with a characteristic green flame.

Copper and Oxygen Compounds.

Red Oxide, Cu_2O and Black Oxide or Copper Monoxide, CuO.

Cuprous or Red Oxide of Copper: Formula—Cu_2O.

This oxide occurs native in ruby-red octahedral crystals. It is artificially prepared by heating equivalent quantities of cupric oxide and copper filings, or by boiling a solution of copper sulphate and sugar, to which excess of caustic potash has been added, the sugar reduces the copper salt, and cuprous oxide is precipitated as a bright red powder. Cuprous oxide imparts to glass a splendid ruby-red color, it forms colorless salts with acids, which rapidly absorb oxygen from the air, and pass into the corresponding cupric compounds. The most important of these salts is cuprous chloride, Cu_2Cl_2, a white solid substance obtained by dissolving a mixture of cupric oxide and metallic copper in hydrochloric acid. The solution of cuprous

chloride possesses the remarkable property of absorbing carbonic oxide gas.

Black Oxide of Copper: Formula—CuO.

This oxide is formed when copper is heated in the air, or when copper nitrate is heated to redness, it yields the blue and green cupric salts, and it is largely used in the laboratory as a means of giving oxygen for the combustion of organic substances. Hydrated copper oxide is obtained as a light blue precipitate when caustic alkali is added to a cupric salt, when this is heated to 100°, it loses its water, and the anhydrous oxide falls as a brown powder. Cupric oxide is soluble in acids, furnishing a series of well crystallizing salts.

Copper, Sulphur and Oxygen Compounds.

Sulphate of Copper: Formula—$CuSO_4$.

This salt is sometimes known as blue vitriol, and is largely manufactured by dissolving copper oxide in sulphuric acid. It crystallizes in large blue crystals. When heated to redness, it loses all its water of crystallization, and forms a white powder, which again at a higher temperature decomposes, leaving copper oxide. Copper sulphate is employed in calico-printing, and in the manufacture of Scheele's and Brunswick green, and other copper pigments.

Copper and Sulphur Compounds.

Sulphide of Copper: Formula—CuS.

The insoluble copper salt, Copper sulphide, is obtained as a black precipitate, when sulphuretted hy-

drogen gas is passed through an acidified solution of a copper salt.

FLUORINE.

Symbol—F. **Atomic Weight—18.9.**

Fluorine is a light greenish-yellow gas, paler and more purely yellow in color than chlorine, possessing a penetrating odor resembling that of hypochlorous acid.

This element occurs combined with the metal calcium, forming calcium fluoride, or fluorspar, a mineral crystallizing in cubes. It also exists in large quantities in cryolite, a mineral found in Greenland, whilst it has been detected in minute quantities in the teeth, and even in the blood, of animals. Fluorine is remarkable as forming no compounds with oxygen, and as being extremely difficult to prepare in a pure state. Many attempts have been made to obtain fluorine in the free state, but none of the methods which yield chlorine, bromine, or iodine give any result. By the action, however, of dry iodine upon dry silver fluoride, it appears that fluorine has been isolated, and it is found to be a colorless gas which does not act upon glass, and is absorbed by caustic potash with the formation of potassium fluoride and hydrogen dioxide.

Fluorine is the most active element with which we are acquainted. It combines explosively with hydrogen in the dark, the direct combination may be shown by simply inverting a jar filled with hydrogen over the positive exit tube of an electrolytic apparatus. As soon as the fluorine comes in contact with the hydrogen a blue, red-bordered flame appears at the end of the platinum tube, hydrofluoric acid being formed

which slowly attacks the glass jar. It decomposes water with the utmost avidity, uniting with the hydrogen to form hydrofluoric acid, and liberating ozone, and therefore in all experiments with the gas the presence of moisture must be excluded as completely as possible.

GOLD.

Symbol—Au. **Atomic Weight—195.7.**

Gold is always found in the metallic state, it occurs in veins in the older sedimentary or in the plutonic rocks, and in the detritus of such rocks, it occurs in traces in the sand of most rivers, and although found generally in small quantities, it is a widely diffused metal. Previous to the discoveries of the gold-fields of California and Australia, it was obtained from certain iron pyrites. In order to obtain the gold, the detritus or sand which contains the metal is washed in a cradle or other arrangement, by means of which the lighter particles of mud or mineral are washed away, whilst the heavier grains of gold sink to the bottom of the vessel. When gold has to be worked in the solid rock, the mineral is crushed to powder, and then shaken up with mercury, and the gold thus extracted by amalgamation.

Gold possesses a brilliant yellow color, and, in thin films, transmits green light, it is nearly as soft as lead. It can be drawn out into fine wire, and is the most malleable of all the metals. It does not tarnish at any temperature, in dry or moist air, nor is it affected by sulphur, like silver, it is not acted upon by any single acid except selenic, but dissolves in presence of free chlorine and in nitro-hydrochloric acid. At high tem-

perature gold is slightly volatile. Pure gold is best prepared by dissolving the ordinary metal in aqua regia, and adding ferrous sulphate, which is oxidized to ferric salt and precipitates the gold as a brown powder.

Gold and Oxygen Compounds.

Gold suboxide: Formula—Au_2O.

Gold trioxide: Formula—Au_2O_3.

Gold unites with oxygen in two proportions, forming Gold suboxide, and Gold trioxide. Neither of these oxides forms salts with acids, but the latter unites with bases to form compounds called aurates, thus potassium aurate is $KAuO_2$. Gold trioxide is obtained by adding zinc oxide or magnesia to a solution of gold chloride, the oxide falls as a brown powder, from which the zinc can be separated by nitric acid. Gold trioxide decomposes, in direct sunlight, into metal and oxygen, and is also reduced when heated to a temperature of about 250°. The most important compound of gold trioxide is fulminating gold. This substance is obtained by acting on a solution of gold with excess of ammonia, a yellow-brown powder is precipitated, which, when dry, explodes very easily when heated to 100°, or when struck with a hammer.

HYDROGEN.

Symbol—H. Atomic Weight—1. Density—1.

Hydrogen is a colorless invisible gas, possessing neither taste nor smell, it is the lightest body known, being 14.47 times lighter than air. It occurs free in small proportions in certain volcanic gases, and it has been lately shown to exist absorbed in certain speci-

mens of meteoric iron, but it is found in much larger quantities, combined with oxygen to form water, and it is by the decomposition of water, or of some other similar hydrogen compound, that the gas is always prepared. One-ninth of the weight of water consists of hydrogen, and this gas can readily be obtained from it by the action of certain metals, which decompose the water, combining with the oxygen to form a metallic oxide, and liberating the hydrogen as a gas. The metals of the alkalies, potassium and sodium, decompose water at the ordinary temperature of the air, some other metals, as iron, are only able to do so at a red heat, whilst others, for instance silver and gold, are unable to decompose water at all. When a small piece of potassium is thrown into water, an instantaneous decomposition of the water ensues, potassium hydroxide (caustic potash) is formed, and the hydrogen of the water is liberated, so much heat being at the same time evolved that the hydrogen takes fire and burns. If the potassium, or, still better, sodium, be wrapped in a piece of wire gauze, as shown in Fig. 12, and thus held in the water of the pneumatic trough, under the mouth of a cylinder, the hydrogen gas thus liberated may be collected, and its properties examined. Water consists of 2 parts by weight of hydrogen and approximately 16 parts by weight of oxygen, and its chemical symbol is therefore H_2O. When potassium or sodium act upon water, half the hydrogen is liberated, the metal taking its place, or water and potassium yield potassium hydroxide and hydrogen. This shows that for every 1 part by weight of hydrogen which is liberated, 39.1 parts by weight of potassium enter into combination. The hydroxide which is

formed dissolves in the water, but its presence can easily be detected either by the peculiar caustic taste which the solution possesses, whence its name, caustic potash, or by its power of turning to a blue color a solution of litmus which has been reddened by an acid.

To prepare hydrogen by the action of red-hot iron on water, a wrought-iron pipe, like a gun-barrel, filled

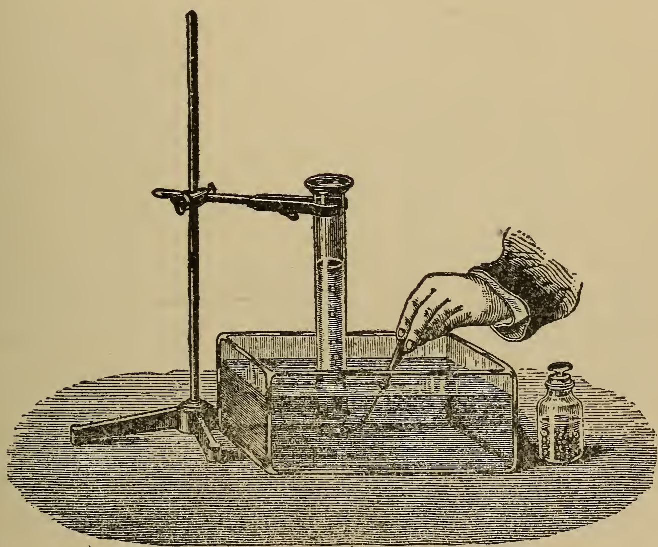

Fig. 12.

with iron turnings, must be heated in a furnace, and steam from a small flask or boiler passed over the red-hot metal through the tube; hydrogen gas is given off, and oxide of iron left in the tube. The most convenient process of preparing pure hydrogen in quantity depends upon a property possessed by those metals, such as iron or zinc, which decompose water at

a red heat, namely, that these metals are able to evolve hydrogen from water at the ordinary temperature of the air if a dilute acid be present. For the purpose of thus obtaining hydrogen, a flask or bottle is provided with a cork and tube as represented in Fig. 13, some zinc clippings are introduced, and a mixture of

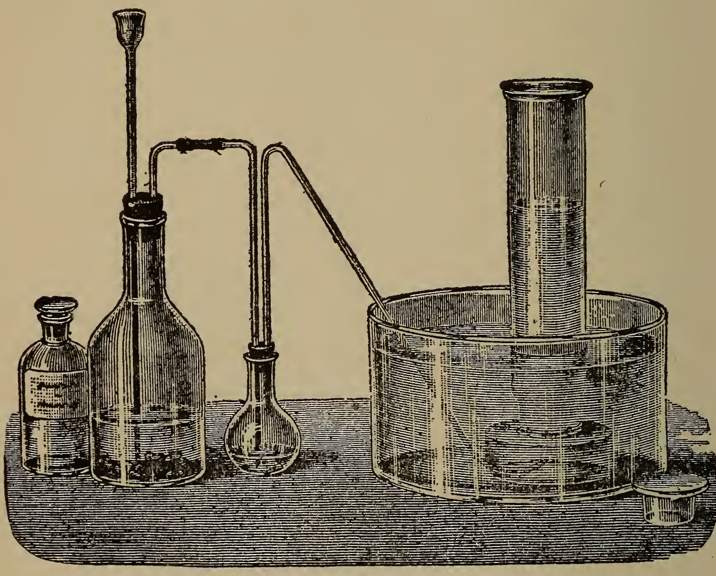

Fig. 13.

one part of sulphuric acid (a compound of sulphur, oxygen, and hydrogen) and eight parts of water poured in through the tube funnel. After a few minutes a rapid effervescence commences, and the evolved gas is collected over water in bottles or cylinders as in the case of oxygen. Care must, however, be taken that all the air is expelled from the flask before the hydrogen

is collected, this is easily ascertained to be the case by filling a test tube with the gas, and trying whether the gas burns quietly when a lighted candle is brought to the mouth of the tube held downwards.

If we concentrate by boiling the liquid remaining in the flask after the evolution of the hydrogen, we find that white crystals separate out when the liquid cools, these consist of zinc sulphate. A given weight of zinc with sulphuric acid and water, can always be made to produce a certain weight of hydrogen, and a certain weight of zinc sulphate will always be formed. It is found by experiment that 2 parts by weight of hydrogen can be obtained by dissolving 65.2 parts of zinc with the formation of 161.2 parts of zinc sulphate. This can be represented by the equation—

$$H_2SO_4+Zn=ZnSO_4+H_2,$$

which not only indicates that sulphuric acid and zinc yield zinc sulphate and hydrogen, but also informs us as to the weights of the respective substances taking part in the reaction.

Hydrogen burns in the air when a light is brought to it with a very slightly luminous, although extremely hot flame, and in the process the hydrogen combines with the oxygen of the air, forming water. The production of water by the combustion of hydrogen in the air may easily be shown by bringing a bright dry glass over the flame of hydrogen issuing from a fine jet, the glass becomes at once dimmed owing to the condensation of water in small drops upon the cold dry surface. A number of these drops can be collected, and, upon examination, they are found to consist of pure water. Hydrogen does not support the combustion of a candle, nor the life of an animal. If a burning taper

is pushed up into a cylinder of this gas, held with its mouth downwards, the hydrogen burns at the mouth of the jar, while the taper is extinguished, it can, however, be relit by the flame at the mouth. Hydrogen can be poured from one vessel to another in the air, but as it is lighter than air it must be poured upwards. The specific gravity of hydrogen, when air is taken as the unit, is found to be 0.0693, but for several reasons it is found to be more convenient to take hydrogen itself as the unit, and compare the weight of the same volumes of other gases with hydrogen instead of air. One litre of hydrogen gas at 0° Centigrade and 760 mm. pressure weighs 0.08936 gram. Free hydrogen, like oxygen, has never been obtained in the liquid or solid state.

Hydrogen and Chlorine Compounds.

Hydrochloric or Muriatic Acid, HCl.

Hydrochloric Acid: Formula—HCl. Molecular Weight—36.2.

This substance, the only known compound of chlorine and hydrogen, is obtained when equal volumes of chlorine and hydrogen are mixed and exposed to the diffused light of day, the gases then combine, and form an unaltered volume of hydrochloric acid gas. If the light be strong, this combination takes place so rapidly that a violent explosion occurs, owing to the sudden disengagement of heat consequent upon this combination. The volume of hydrochloric acid formed is equal to that of the chlorine and hydrogen, one molecule of hydrogen and one molecule of chlorine give two molecules of hydrochloric acid.

Hydrochloric acid may, however, be more easily prepared by heating common salt (sodium chloride) and sulphuric acid in a flask.

Sodium chloride and sulphuric acid give hydrochloric acid and hydrogen sodium sulphate.

Hydrochloric acid is a colorless gas, 1.269 times heavier than air, it fumes strongly in damp air, combining with the moisture, and has a strongly acid reaction. It is very soluble in water, one volume of this liquid at 15° dissolving 454 volumes of the gas. This solution is the ordinary hydrochloric or muriatic acid of commerce. Under a pressure of 40 atmospheres the gas forms a limpid liquid. The gas can be collected over mercury, and its solubility in water strikingly shown by allowing a few drops of water to ascend to the surface of the mercury in contact with the gas. A rapid rise of the mercury in the jar immediately occurs. It fumes strongly in the air, and, when heated in a retort, loses at first hydrochloric acid gas, but after a time an aqueous acid distils over at the ordinary atmospheric pressure, containing 20.22 per cent of HCl, and boiling constantly at 110°. If the distillation be conducted under a diminished pressure, the acid boils constantly at a lower temperature, and attains a composition which is different for each boiling point, hence the constant acids thus obtained by boiling the solution of hydrochloric acid gas in water cannot be considered as definite compounds of HCl and water. This fact holds good for many other aqueous solutions of acids, etc., that residues constantly boiling at the same temperature, and having constant compositions, are obtained on distillation, the composition

and boiling point varying, however, with the pressure under which the distillation has been conducted.

Enormous quantities of hydrogen chloride, commonly called muriatic acid, from muria, sea salt, are obtained as a by-product in the manufacture of sodium carbonate. The acid thus produced is, however, very impure, having a yellow color, and containing iron, arsenic, organic matter, and sulphuric acid in solution.

The exact composition of hydrogen chloride is best determined by decomposing the aqueous acid in the dark by means of a current of electricity, and collecting the gases (hydrogen and chlorine) evolved in a long tube, after allowing the decomposition to go on for some time. If the tube thus filled be opened in the dark under a solution of potassium iodide, the solution will rise in the tube, the iodine being liberated, the chlorine combining with the potassium, until exactly half the tube is filled with liquid, the remaining gas is found to consist of hydrogen. If the mixture of electrolytic gases, which can with care be sealed up in a strong tube having very finely drawn out ends, be exposed to the action of daylight, or of a bright artificial light, such as that of burning magnesium wire, immediate combination of the two gases will ensue, and on opening one of the ends under water this liquid will completely fill the whole of the tube, showing that the component gases were present in exactly the proportion needed to form hydrochloric acid gas which dissolves in the water.

Aqua Regia. Certain metals, such as gold and platinum, and many metallic compounds, such as certain sulphides, which do not dissolve in either nitric or hydrochloric acid separately, are readily soluble in a

mixture of both of these acids, especially upon warming. This mixture has been termed Aqua Regia, and its solvent action depends upon the fact that it contains free chlorine, liberated by the oxidizing action of nitric acid on the hydrogen of the hydrochloric acid. The metals combine directly with this free chlorine to form soluble chlorides, and the sulphides are decomposed by it. The nitric acid is reduced to nitrogen dioxide, and this combines with a portion of the chlorine to form the compounds $NOCl$ and $NOCl_2$, which are liberated as yellowish gases, condensing to a dark yellow, very volatile liquid when they are led into a freezing mixture. The same compounds are formed by direct combination when two gases, nitrogen dioxide and chlorine, are mixed together. If chlorine is present in excess, $NOCl_2$ is formed, while $NOCl$ is produced when the oxide of nitrogen is present in largest quantity.

Hydrogen and Fluorine Compounds.

Hydrofluoric Acid.

Symbol—HF. **Molecular Weight—19.9.**

Hydrofluoric acid gas must be prepared in a leaden or platinum vessel, as glass is rapidly attacked by the vapor. The colorless gas thus obtained fumes strongly in the air, if it be passed into a metallic tube placed in a freezing mixture at the temperature of —20°, a liquid is formed, this liquid is strong aqueous hydrofluoric acid. It appears doubtful whether the acid has been obtained in the liquid state. The strong acid acts very violently upon the skin, producing painful

wounds, and the fumes of the gas are likewise dangerous from their corrosive power. When brought into contact with water the strong acid dissolves with a hissing noise. This aqueous acid attains a constant boiling point under the ordinary atmospheric pressure, when the liquid contains 37 per cent.

The most remarkable property of hydrofluoric acid is its power of etching upon glass. This arises from the fact that fluorine forms, with the silicon contained in the glass.

The acid thus obtained is a highly dangerous substance, and requires the most extreme care in its manipulation, the inhalation of its vapor having produced fatal effects. A drop on the skin gives rise to blisters and sores which only heal after a very long period. From its great volatility the anhydrous acid can only be safely preserved in platinum bottles having a flanged mouth, a platinum plate coated with paraffin being tightly secured to the flanged mouth by clamp screws. The acid must be kept in a cool place not above a temperature of 15°, otherwise it is very likely to burst the bottle, and a freezing mixture should always be at hand when experimenting with it.

Anhydrous hydrofluoric acid can also be obtained by acting on dry silver fluoride with hydrogen.

If the hydrofluoric acid is not required to be perfectly anhydrous a much easier process than the foregoing can be adopted. This consists in the decomposition of fluor-spar by strong sulphuric acid, when hydrofluoric acid and calcium sulphate are formed.

For this preparation vessels of platinum, or, on a large scale, vessels of lead, can be employed.

Hydrogen and Oxygen Compounds.

Water or Hydrogen Monoxide, H_2O and Peroxide of Hydrogen or Hydrogen Dioxide, H_2O_2.

Water: Formula—H_2O. Molecular Weight—17.88. Density—8.95.

When hydrogen is burnt in the air, water is formed by the union of the hydrogen gas with the oxygen of the air. Cavendish first ascertained in 1871, that by the combustion of two volumes of hydrogen and one

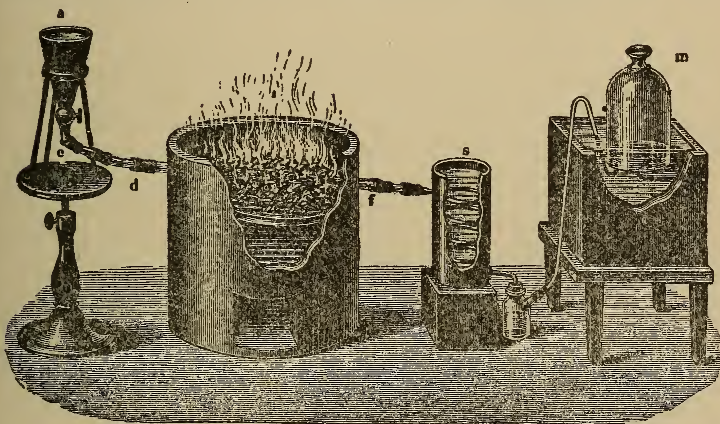

Fig. 14.

of oxygen, pure water is formed and nothing else. Lavoisier in 1783, the French chemist, repeated and confirmed the experiments of Cavendish. The apparatus used by him for proving that hydrogen gas is really contained in water, is seen in facsimile in Fig. 14. The water contained in the vessel **a** was allowed to drop slowly into the tube, **e d**, from which it flowed into the gunbarrel, **d f**, heated to redness in the fur-

nace. Here part of the water was decomposed, the oxygen entering into combination with the metallic iron, whilst the hydrogen and some undecomposed steam passed through the worm, **s**, where the steam was condensed and the hydrogen was collected and measured in the glass bell-jar, **m**. The result of these experiments was found to be that 13.13 parts by weight of hydrogen united to 86.87 parts by weight of oxygen, or 12 volumes of oxygen with 22.9 volumes of hydrogen.

Bunsen's modification of Cavendish's method consisted in bringing known volumes of the two gases successively into a eudiometer, and combining them by means of an electric spark, carefully observing the consequent change in volume. The eudiometer is a strong glass tube **c**, Fig. 15, one metre in length and 2.5 centimeters in diameter, closed at the top, open at the bottom, and having platinum wires sealed through the glass near the closed end. The tube is accurately graduated with a millimeter scale. The eudiometer contains at the top, one drop of water to render the gases moist. It is first completely filled with mercury and then inverted in the trough **d**, containing the same metal.

Hydrogen gas is now allowed to enter the tube, and the volume admitted measured, equal to 100 volumes, oxygen gas is next admitted, and the volume of the two mixed gases measured, 75 volumes of oxygen are added. In making this experiment, care must, however, be taken that the temperature and atmospheric pressure are carefully measured by means of the thermometer **b** and the barometer **a**, shown in the figure, it is also necessary that the tube be not more than half

full of the gaseous mixture, as great heat is evolved by the combustion, and hence a sudden expansion of

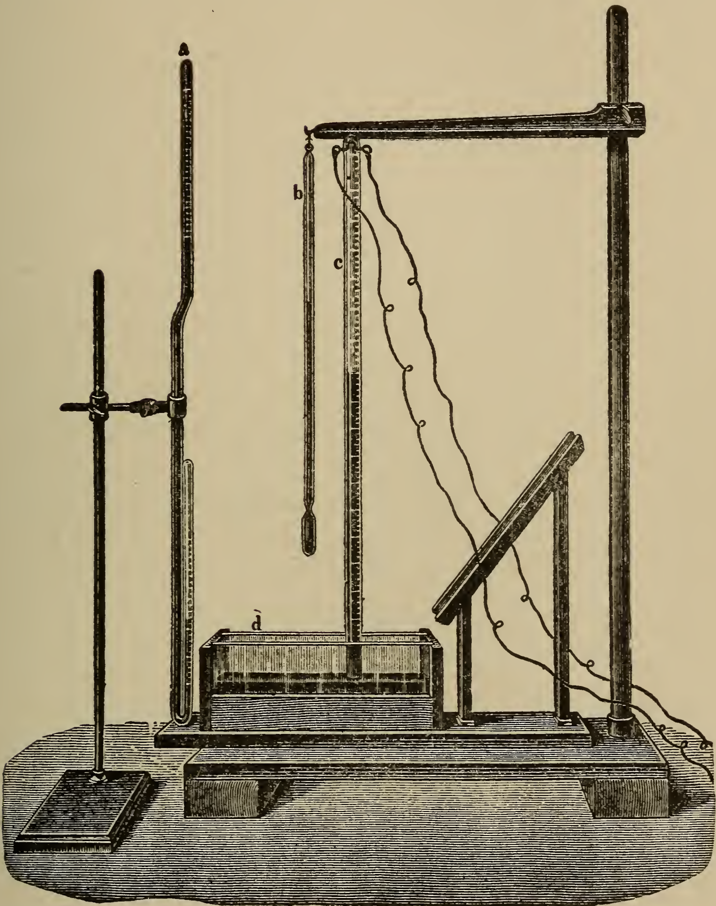

Fig. 15.

volume occurs, for which reason it is necessary to press down the open end of the tube upon a plate of soft rubber placed under the mercury. An electric spark

is now passed through the gas along the platinum wires, when a flame is seen to pass down through the gas, showing that combination has occurred, the water produced will be deposited as dew upon the inside of the tube, and will then only take up about $\frac{1}{2000}$ part of the bulk which its constituent gases occupied, so that its volume may be neglected. When the bottom of the eudiometer is opened, the column of mercury in the tube rises, and then only 25 volumes of gas remain, and this turns out to be pure oxygen. Thus 100 volumes of hydrogen require exactly 50 volumes of oxygen for their complete combustion. By a modification of this experiment, it can be shown that the volume of the gaseous water formed occupies exactly 100 volumes, or 2 volumes of hydrogen unite with 1 of oxygen to form 2 volumes of steam, hence the density of steam or weight of 1 volume is $\frac{16+2}{2}=9$. The most striking method of demonstrating the composition of water analytically is by splitting it up into its constituent gases by means of a current of electricity. For this purpose fill a glass vessel as in Fig. 16, with water acidulated with sulphuric acid to enable it to conduct the electricity, and bring two test tubes filled with water and inverted into this vessel over two small platinum plates attached to wires of the same metal passing through the soft rubber stopper at the bottom of the glass bowl, on connecting these with the terminals of a battery of three or four elements, an evolution of gas from each plate is noticed, that disengaged from the plate in connection with the carbon pole of the battery is found to be pure oxygen, whilst that coming off from the other plate connected with

the zinc pole of the battery, is pure hydrogen gas. If the two tubes be graduated, it will be seen that the volume of the hydrogen is a very little more than

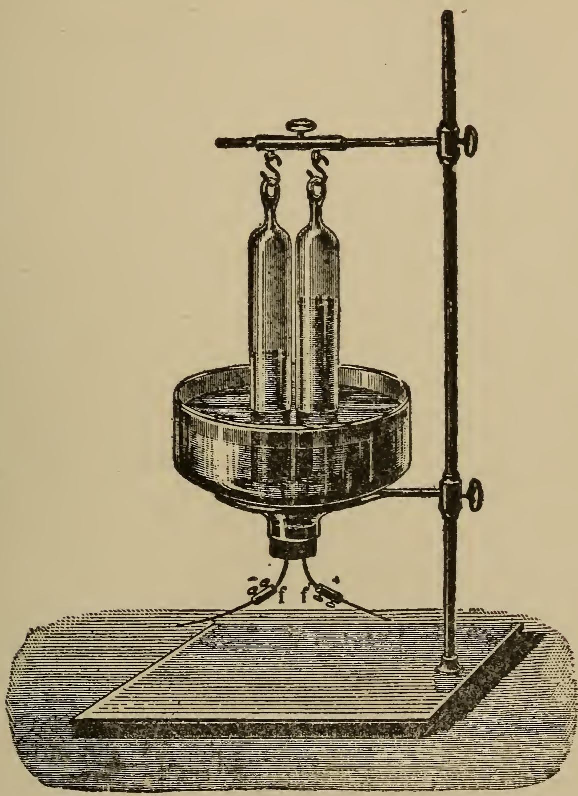

Fig. 16.

double that of the oxygen, owing to the oxygen being rather more soluble in water than the hydrogen.

The water formed by the combination of the two gases can again be decomposed into free hydrogen and oxygen, and this can readily be effected by a current

of electricity. For the purpose of exhibiting this pass a current of electricity from four or six elements by means of two platinum poles through some water acidulated with sulphuric acid as in Fig. 16. The instant contact is made, bubbles of gas begin to ascend from each platinum plate and collect in the graduated tubes, which at first are filled with the acidulated water. After a little time it will be seen that the plate which is in connection with the zinc of the battery evolves more gas than the one which is in contact with the carbon of the battery, and after the evolution has continued for a few minutes one tube will be seen to contain twice as much gas as the other. On examination, the larger volume of gas will be found to be hydrogen, and will take fire and burn when a light is brought to the end of the tube in which it was collected, whilst the smaller volume of gas is seen to be oxygen, a glowing chip of wood being rekindled when plunged into it.

Whereas the combination of hydrogen with oxygen is attended by the evolution of energy in the forms of light and heat, the inverse change, the decomposition of water, requires energy to be supplied, and in the experiment just described, this is conveyed to the water in the form of the energy of the electrical current.

In order to collect the mixed explosive gases evolved by the electrolytic decomposition of water, an apparatus such as is shown in Fig. 17 is used.

Oxygen being 16 (exactly 15.88, hydrogen as 1) times as heavy as hydrogen, and these gases combining to form water in the proportions by volume of one volume of the former to two of the latter, we now know that the proportions by weight in which these

gases exist in water must be as 16 to 2. It is nevertheless most important that this calculation be verified by direct experiment. For this purpose, use is made of the fact that copper oxide when heated alone does not part with any of its oxygen, but when heated in

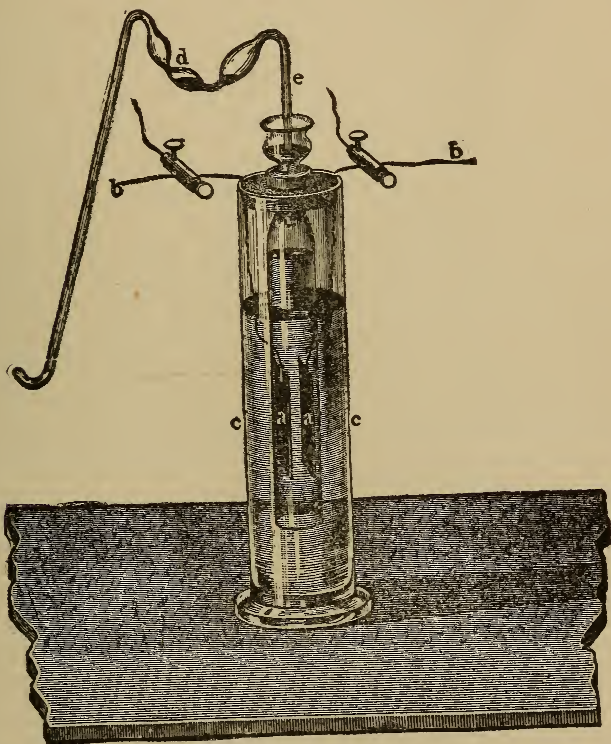

Fig. 17.

presence of hydrogen it parts with as much oxygen as will, by combining with the hydrogen, form water, being itself wholly or partly reduced to metallic copper. If, therefore, a known weight of copper oxide

is taken, heat and pure hydrogen passed over it until it has parted with all its oxygen, and if we collect and weigh all the water thus formed, and likewise weigh the remaining metallic copper, we shall have made a synthesis by weight of water. For the loss in weight of the copper oxide is the weight of oxygen which has combined with hydrogen to form water, and the difference between this weight and that of the water formed, is the weight of the hydrogen thus combined. The arrangement used for this determination is represented in Fig. 18. The hydrogen evolved by zinc from sulphuric acid in the bottle on the left hand is purified from any trace of arsenic, sulphur, and moisture which it may contain by passing through the U tubes, containing absorbent substances. The tube containing a very hygroscopic substance, is weighed both before and after the experiment, and if no increase occurs, the dryness of the gas is insured. The gas then comes in a perfectly pure state into contact with the heated copper oxide contained in the bulb B. This first bulb, which is accurately weighed, is placed in connection with a second bulb B, in which the water formed by the reduction of the oxide collects, any moisture which may escape condensation in this bulb is retained in weighed drying tubes, containing fragments of pumice stone moistened with sulphuric acid. Most careful experiments made according to this method, carried out with many precautions which cannot here be detailed, have shown that 88.89 parts of oxygen by weight unite with 11.11 parts of hydrogen to form 100 parts of water.

Free oxygen and hydrogen combine together, when a light is brought in contact with them, with so much

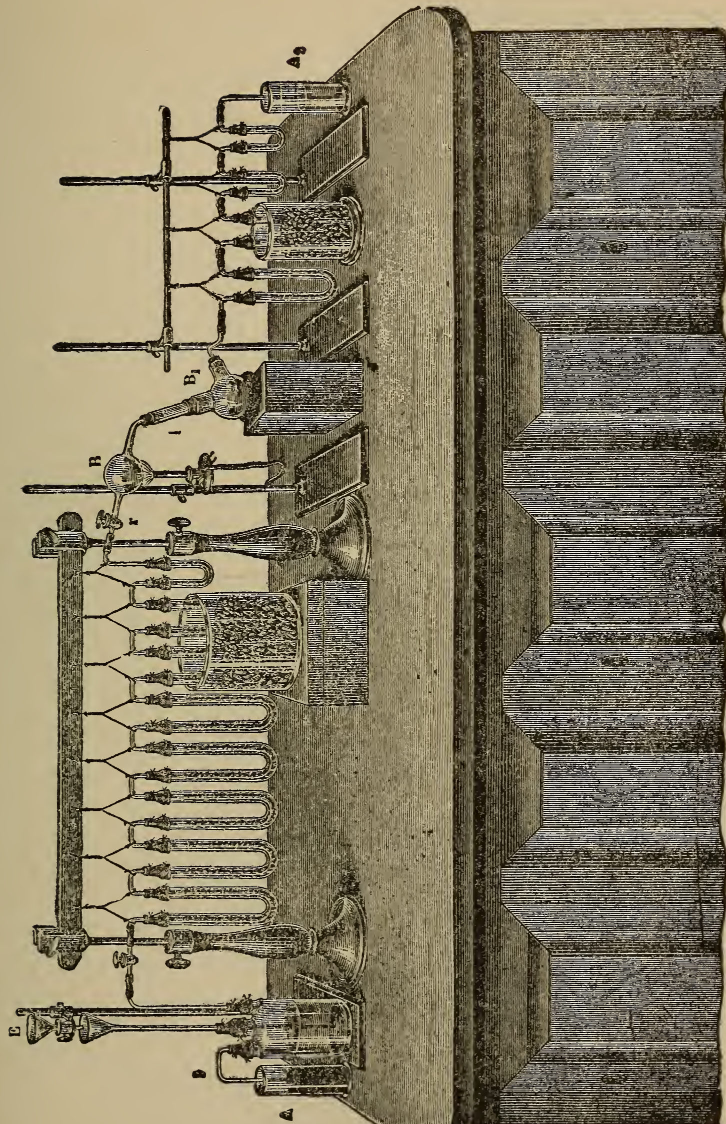

Fig. 18.

force that a violent and dangerous explosion occurs from the sudden expansion caused by the great heat evolved in combination. If we fill a strong soda-water bottle one-third full with oxygen and two-thirds with hydrogen, and then bring a flame to the mouth, the gases combine, producing a sudden detonation like the report of a pistol. Many fatal accidents have occurred to persons who have carelessly experimented with large volumes of this explosive mixture. In order to exhibit the great heat evolved by the combination of the two gases, the oxyhydrogen blowpipe is employed, in this arrangement the gases are contained separately in two rubber bags, being only brought together at the point at which the combination is desired, so that all danger of explosion is avoided. The flame thus produced is very slightly luminous, but its temperature is so high, that the most difficultly fusible metals, such as platinum, may be easily melted in it, whilst iron wire held in the flame burns with beautiful scintillations, forming an oxide of iron. A piece of chalk or lime placed in this flame becomes heated to bright whiteness and emits an intense light, often used for signal purposes.

Water exists in nature in three forms: In the solid form as ice, in the liquid state as water, and in the gaseous form as steam. At all temperatures between 0° and 100° Centigrade it takes the liquid form, and above 100° it entirely assumes the gaseous form, under the ordinary atmospheric pressure of 760 mm. The melting point of ice is always found to be a constant temperature, and hence it is taken as the zero of the Centigrade scale, water may, however, under certain conditions, be cooled below 0° Centigrade without be-

coming solid, still ice can never exist at a temperature above 0° Centigrade. In passing from the solid to the liquid state water becomes reduced in volume, and on freezing a sudden expansion, from 1 volume to 1.099 takes place. That this expansion exerts an almost irresistible force is well illustrated by the splitting of rocks during the winter. Water penetrates into the cracks and crevices of the rocks, and on freezing widens these openings, this process being repeated over and over again, the rock is ultimately split into fragments. Hollow balls of thick castiron can thus easily be split in two by filling them with water and closing by a tightly-fitting screw, and then exposing them to a temperature below 0° Centigrade.

In the passing from solid ice to liquid water, we not only observe this alteration in bulk, but we notice that a very remarkable absorption, or disappearance of heat, occurs. This is rendered plain by the following simple experiment: Take a kilogram of water at the temperature 0°, and another kilogram of water at 79°, if we mix these, the temperature of the mixture will be the mean, or 39°.5, if, however, we take 1 kilogram of ice at 0° and mix it with a kilogram of water at 79°, we shall find that the whole of the ice is melted, but that the temperature of the resulting 2 kilograms of water is exactly 0°, in other words, the whole of the heat contained in the hot water has just sufficed to melt the ice, but has not raised the temperature of the water thus produced. Hence we see that in passing from the solid to the liquid state a given weight of water takes up or renders latent just so much heat as would suffice to raise the temperature of the same weight of water through 79° Centigrade, the latent

heat of water is therefore said to be 79 thermal units —a thermal unit meaning the amount of heat required to raise a unit weight of water through 1° Centigrade. When water freezes, or becomes solid, this amount of heat which is necessary to keep the water in the liquid form, and is therefore well termed the heat of liquidity, is evolved, or rendered sensible. A similar disappearance of heat on passing from the solid to the liquid state, and a similar evolution of heat on passing from the liquid to the solid form, occurs with all substances, the amount of heat thus rendered latent or evolved varies, however, with the nature of the substance. A simple means of showing that heat is evolved on solidification consists in obtaining a saturated hot solution of Glauber's salt, sodium sulphate, and allowing it to cool. Whilst it remains undisturbed, it retains the liquid form, but if agitated it at once begins to crystallize, and in a few moments becomes a solid mass. If a delicate thermometer be now plunged into the salt while solidifying, a sudden rise of temperature will be noticed. Similarly water at rest may be cooled down below 0° Centigrade without solidifying, but if agitated it at once solidifies, and the temperature of the whole mass instantly rises to 0° Centigrade.

When water is heated from 0° to 4°, it is found to contract, thus forming a striking exception to the general law, that bodies expand when heated and contract on cooling, on cooling from 4° to 0° it expands again. Above 4°, however, it follows this ordinary law, expanding when heated, and contracting when cooled. This peculiarity in the expansion and contraction of water may be expressed by saying that the point of maximum density of water is 4° Centigrade,

that is, a given bulk of water will at this temperature weigh more than at any other. Although the amount of contraction on heating from 0° to 4° is but small, 1 volume of water at 4° becoming 1+0.00012 at 0°, it yet exerts a most important influence upon the economy of nature. In order better to understand what the state of things would be if water obeyed the ordinary laws of expansion by heat, we may perform the following experiment: Take a jar containing water at a temperature above 4°, place a thermometer at the top and another at the bottom of the liquid. Now bring the jar into a place where the temperature is below the freezing point, and observe the temperature at the top and bottom of the liquid as it cools. It will be seen that at first the upper thermometer always indicates a higher temperature than the lower one, after a short time both thermometers mark 4°, and, as the water cools still further, it will be seen that the thermometer at the top always indicates a lower temperature than that shown by the one at the bottom. This cooling goes on till the temperature of the top layer of water sinks to 0°, after which a crust of ice is formed, but if the mass of the water be sufficiently large, the temperature of the water at the bottom is never reduced below 4°. In nature precisely the same phenomenon occurs in the freezing of lakes and rivers, the surface-water is gradually cooled by cold winds, and thus becoming heavier, sinks, whilst lighter and warmer water rises to supply its place, this goes on till the temperature of the whole mass is reduced to 4°, after which the surface-water never sinks, however much it be cooled, as it is always lighter than the deeper water at 4°. Hence ice is formed only at the

top, the mass of water retaining the temperature of 4°. Had water become heavier as it cooled down to the freezing point, a continual circulation would be kept up until the whole mass was cooled to 0°, when solidification of the whole would ensue. Thus our lakes and rivers would be converted into solid masses of ice, which the summer's warmth would be quite insufficient thoroughly to melt; hence the climate of our now temperate zone might approach in severity that of the Arctic regions. Sea-water does not freeze en masse, owing to the great depth of the ocean, which prevents the whole from ever being cooled down to the freezing point, similarly very deep lakes never freeze, as the temperature of the whole mass never gets reduced to 4° Centigrade.

In passing from the liquid to the gaseous state, water exhibits several interesting and important phenomena. In the first place, when we heat water to 100° C. it begins to boil, or enters into ebullition, that is, a rapid disengagement of water-gas, or steam, from the lower or most heated surface takes place, this is well seen when water is heated in a glass globe over a gas flame. In this passage from the liquid to the gaseous state, a large quantity of heat becomes latent, the temperature of the steam given off being the same as that of the boiling water, as, like all other bodies, water requires more heat for its existence as a gas than as a liquid. The amount of heat latent in steam is roughly ascertained by the following experiment. . Into 1 kilogram of water at 0°, steam from boiling water, having the temperature of 100°, is passed until the water boils, it is then found that the whole weighs 1.187 kilos., or 0.187 kilo. of water in the form of

steam at 100° has raised 1 kilo. of water from 0° to 100°; or 1 kilo. of steam at 100° would raise 5.36 kilos. of ice-cold water through 100°, or 536 kilos. through 1°. Hence the latent heat of steam is said to be 536 thermal units.

Whenever water evaporates or passes into the gaseous state, heat is absorbed, and so much heat may be thus abstracted from water that it may be made to freeze by its own evaporation. A beautiful illustration of this is found in an instrument consisting of a bent tube, having a bulb on each end, and containing water and vapor of water, but no air. On placing all the water in one bulb, and plunging the empty bulb into a freezing mixture, a condensation of the vapor of water in this empty bulb occurs, and a corresponding quantity of water evaporates from the other bulb to supply the place of the condensed vapor, this condensation and evaporation go on so rapidly that in a short time the water cools down below 0°, and a solid mass of ice is left in the bulb. By a very ingenious arrangement this plan of freezing water by its own evaporation has been practically carried out on a large scale by M. Carre, by means of which ice can be most easily and cheaply prepared. This arrangement consists simply of a powerful air-pump A, Fig. 19, and a reservoir B of a hygroscopic substance, such as strong sulphuric acid. On placing a bottle of water in connection with this apparatus, and on pumping for a few minutes, the water begins to boil rapidly, and the temperature of the water is cooled so low by its own evaporation as to freeze to a mass of ice.

Water, and even ice, constantly give off steam or aqueous vapor at all temperatures, when exposed to

the air, thus we know that if a glass of water be left in a room for some days, the whole of the water will gradually evaporate. This power of water to rise in vapor at all temperatures is called the elastic force, or pressure, of aqueous vapor, it may be measured, when a small quantity of water is placed above the mercury in a barometer, by the depression which the

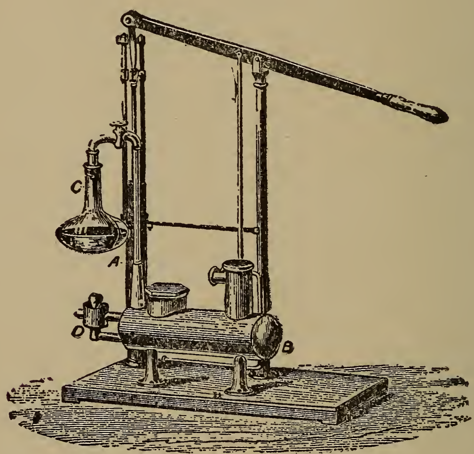

Fig. 19.

tension of the vapor thus given off is capable of exerting upon the mercurial column. If we gradually heat the drops of water thus placed in the barometer, we shall notice that the column of mercury gradually sinks, and when the water is heated to the boiling point, the mercury in the barometer tube is found to stand at the same level as that in the trough, showing that the elastic force of the vapor at that temperature

is equal to the atmospheric pressure. Hence water boils when the pressure of its vapor is equal to the superincumbent atmospheric pressure. On the tops of mountains, where the atmospheric pressure is less than at the sea's level, water boils at a temperature below 100°. At Quito, where the mean height of the barometer is 527 mm., the boiling point of water is 90°.1, that is, the pressure of aqueous vapor at 90°.1 is equal to the pressure exerted by a column of mercury 527 mm. high. Founded on this principle, an instrument has been constructed for determining heights by noticing the temperatures at which water boils. A simple experiment to illustrate this fact consists in boiling water in a globular flask, into the neck of which a stopcock is fitted, as soon as the air is expelled, the stopcock is closed, and the flask removed from the source of heat, the boiling then ceases, but on immersing the flask in cold water, the ebullition recommences briskly, owing to the reduction of the pressure consequent upon the condensation of the steam, the pressure of the vapor at the temperature of the water in the flask being greater than the diminished pressure. All other liquids obey a similar law respecting ebullition, but as the pressure of their vapors are very different, their boiling points vary considerably.

When steam is heated alone, it expands according to the law previously given for permanent gases, but when water is present, and the experiment is performed in a closed vessel, the elastic force of the steam increases in a far more rapid ratio than the increase of temperature. The following table gives the pressure of aqueous vapor, as determined by experiment, at different temperatures measured on the thermometer.

PRESSURE OF THE VAPOR OF WATER.

Temperature Centigrade.	Tension in millimetres of mercury.	Temperature Centigrade.	Pressure in atmospheres, 1 atmosphere =760 mm. of mercury.
−20°	0.927	100°	1.
−10	2.093	111.7	1.5
0	4.600	120.6	2.
+5	6.534	127.8	2.5
10	9.165	133.9	3.
15	12.699	144.0	4.
20	17.391	159.2	6.
30	31.548	170.8	8.
40	54.906	180.3	10.
50	91.982	188.4	12.
60	148.791	195.5	14.
70	233.093	201.9	16.
80	354.280	207.7	18.
90	525.450	213.0	20.
100	760.000	224.7	25.

We now see why the barometric height must be noticed in graduating a thermometer, if the height differ from 760 mm. the temperature of the water boiling under that pressure will not be quite 100° Centigrade. A metal vessel is here employed, because it is found that water does not always boil at 100° in glass vessels, even though the atmospheric pressure be 760 mm., owing to some molecular action analogous to cohesion between the glass and water.

Pure water and ice, when seen in large masses, are found to possess a blue color. In order to obtain pure water the chemist is obliged to distill river or spring-water, that is, to boil the water and collect the water formed by the condensation of the steam thus produced, as all such water contains more or less solid

matter in solution derived from the surface of the earth over which the water flows, this dissolved solid matter is left behind on boiling off the water. Solid matter in suspension can be got rid of by the simpler process of filtration through paper, sand, etc. An arrangement for distillation is seen in Fig. 20, it consists of a retort B in which the impure water is placed, connected with a condenser D, made of tubes, between which a current of cold water is made to circulate.

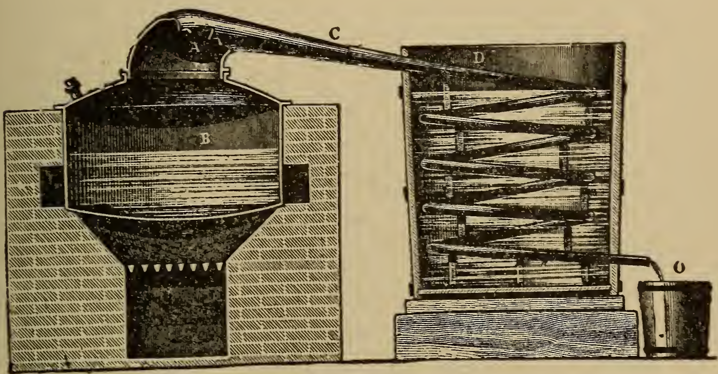

Fig. 20.

The distilled water is collected in the flask O placed at the end of the apparatus. Rain-water is the purest form of water occurring in nature, but even this contains impurities derived from the dust, etc., in the air, and no sooner does it touch the earth's surface than it dissolves some of the materials with which it comes in contact, and according to the nature of the ground over which it passes becomes more or less impure. All fresh-water on the earth's surface has been derived from the ocean by a vast process of distillation, hav-

ing been deposited in the form of rain or snow from the atmosphere.

All the rain-water ultimately passes in the form of spring-water, or river-water, into the sea, carrying with it the soluble constituents which have been dissolved out from the strata through which it has percolated. In consequence of this continual accession of soluble salts, and removal of pure water by evaporation, the sea-water is rendered salt, it contains about 35 parts of solid matter, 28 of which consist of common salt or sodium chloride in solution in 1,000 parts of water.

Water is the most general solvent for chemical substances with which we are acquainted. Most salts are soluble to a greater or less extent in water, and are deposited again in crystals when the water is evaporated, we are unacquainted with any simple general law regulating the quantities of salts taken up by water, in most cases the solubility is greater in hot than in cold water. Water is also contained in the solid state in combination as water of crystallization in many salts, when this water is driven off by heat, the crystal falls to powder. Gases also dissolve in water in quantities varying with the nature of the gas, the temperature, and the pressure to which the gas is subjected. It is solely in consequence of the presence of oxygen derived from the air dissolved in the water of lakes, rivers, and seas, that fish are enabled to keep up their respiration, as the water passes through their gills the oxygen is taken up to purify their blood.

Peroxide of Hydrogen or Hydrogen Dioxide: Formula—H_2O_2. Weight—33.76.

This substance has received the name of oxygenated water, as it easily decomposes into oxygen and water. It is found to contain twice as much oxygen as water does, consisting of 2 parts by weight of hydrogen combined with 32 of oxygen, hence, if we represent water by the symbol H_2O, hydrogen dioxide will be written H_2O_2. It does not occur in nature, but is artificially prepared by acting on barium dioxide, BaO_2, with hydrochloric acid, H_2Cl_2, an exchange takes place between the barium and hydrogen, giving rise to hydrogen dioxide and barium chloride.

Hydrogen dioxide may also be prepared by passing carbonic acid gas through barium dioxide suspended in water, when barium carbonate separates out as a white powder insoluble in water, and hydrogen dioxide remains in solution. The reaction is represented by the following equation:

$$BaO_2+H_2O+CO_2=BaCO_3+H_2O_2.$$

The aqueous solution of the dioxide is concentrated by allowing the water to evaporate under the receiver of an air-pump, the liquid after a time becomes thick, but it cannot be entirely freed from water. Hydrogen dioxide is chiefly characterized by the ease with which it loses half its oxygen, this gas is slowly given off at 20°, but at 100° Centigrade the evolution of oxygen becomes very rapid. In consequence of the readiness with which it gives off oxygen, hydrogen dioxide acts as a powerful bleaching agent, rapidly oxidizing and destroying vegetable coloring matter. A remarka-

ble decomposition occurs when this substance is brought in contact with ozone, common oxygen and water being produced. Another interesting reaction occurs when silver oxide is brought together with hydrogen dioxide, as the silver oxide is reduced to metallic silver whilst water and common oxygen are formed.

Hydrogen and Sulphur Compounds.

Hydrogen Sulphide, H_2S and Hydrogen Disulphide, H_2S_2.

Hydrogen Sulphide or Sulphuretted Hydrogen: Formula—H_2S. Molecular Weight—33.83.

This gas is best prepared by the action of dilute sulphuric acid upon iron sulphide, FeS, iron sulphate being also formed, thus:

$$FeS+H_2SO_4=FeSO_4+H_2S,$$

where two atoms of hydrogen change place with one of divalent iron. Fig. 21 represents a convenient form of apparatus for the production and purification of this gas. It may be collected over warm water, and is a colorless gas, possessing the peculiar odor of rotten eggs, it burns on application of a light with a bluish flame, forming water and sulphur dioxide. When inhaled, it acts as a poison on the animal economy, even if diluted with large quantities of air. Sulphuretted hydrogen gas dissolves in water to a considerable extent, imparting its peculiar smell and a slightly acid reaction to the water. One volume of water at 0° dissolves 4.37 volumes of the gas, whilst at 15° 3.23 volumes are soluble. Exposed to a tempera-

ture of —74°, this gas condenses to a colorless, mobile liquid, which, when further cooled to —85° freezes to a transparent ice-like solid. Under a pressure of about seventeen atmospheres this gas liquefies at the ordinary temperature of the air. Sulphuretted hydrogen occurs free in nature in volcanic gases, as well as in the water of certain springs. Harrogate waters owe their peculiar odor and medicinal power to the presence of this gas. It is likewise generated by the

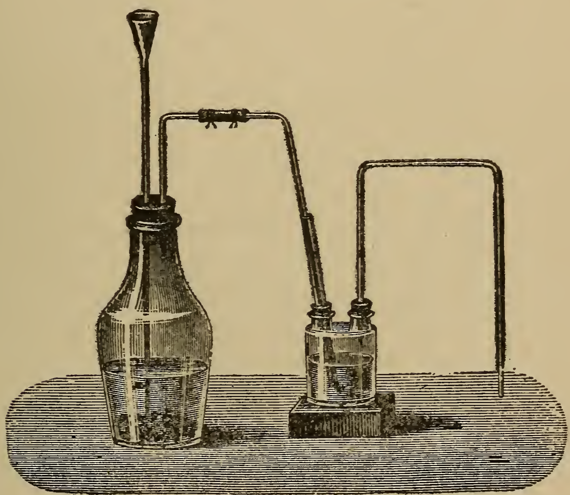

Fig. 21.

putrefaction of animal matters, such as albumen, or the white of eggs, which contains sulphur, also by the deoxidation of sulphates in presence of decaying organic matter.

The composition of sulphuretted hydrogen may be ascertained either by heating a small piece of metallic tin in a known volume of the gas, when tin sulphide

will be formed, and hydrogen liberated, or by decomposing the gas by means of a red-hot platinum wire, when the whole of the sulphur is deposited, and hydrogen set free. In both cases the volume of hydrogen obtained is found to be equal to that of the gas employed, and hence 2 volumes of sulphuretted hydrogen, weighing 34, consist of 2 volumes of hydrogen, weighing 2, and 1 volume of sulphur vapor, weighing 32.

Sulphuretted hydrogen is an invaluable re-agent in the laboratory, as by its means we are enabled to separate the metals in groups. If we pass a current of this gas through a solution of a copper salt, to which a small quantity of acid has been added, we obtain an immediate precipitate of copper sulphide, thus

$$CuSO_4+H_2S=CuS+H_2SO_4.$$

If we do the same with a solution of an iron salt, we get no such precipitate, because iron sulphide is soluble in an acid, but on the addition of an alkali, iron sulphide is at once precipitated, thus

$$FeSO_4+2KHO+H_2S=FeS+K_2SO_4+2H_2O.$$

We may thus divide the metals into groups: First, those which, like copper are precipitated by sulphuretted hydrogen from an acid solution, or the copper group, second, those which are not precipitated by sulphuretted hydrogen in an acid solution, but which are so precipitated in an alkaline one, or the iron group, and third, those which are in no case precipitated by this re-agent, as their sulphides are soluble either in water, acids, or alkalies, to this group belong the metals of the alkalies and alkaline earths.

Hydrogen Disulphide: H_2S_2.

This substance is an oily liquid obtained by pouring a solution of calcium disulphite into hydrochloric acid. This compound is not used in the arts, but it is interesting as having a composition corresponding to H_2O_2, and because, like this substance, it possesses bleaching properties. It easily decomposes into sulphuretted hydrogen and free sulphur.

Hydrogen, Carbon and Nitrogen Compounds.

Hydrocyanic or Prussic Acid: Formula—HCN.

Hydrocyanic acid is generally prepared from potassium ferrocyanide, which is decomposed by dilute sulphuric acid. The process enables us to prepare any desired quantity, either of the aqueous acid of different strengths, or of the anhydrous compound. A cold mixture of 14 parts of water and 7 parts of concentrated sulphuric acid is poured upon 10 parts of coarsely powdered potassium ferrocyanide contained in a large retort the neck of which is placed upwards in a slanting direction. If the anhydrous acid be required, the vapor is allowed to pass through cylinders or U-tubes filled with calcium chloride and surrounded by water heated to 30°, after which the gas, thus dried, is condensed by passing into a receiver surrounded by ice or a freezing mixture. To prepare the aqueous acid the drying apparatus may be omitted and the gas evolved from a flask as shown in Fig. 22, passed through a Liebig's condenser, and then led into the requisite quantity of distilled water. The action

of the dilute sulphuric acid upon potassium ferrocyanide is represented by the following equation

$$2Fe(CN)_6K_4+3H_2SO_4=3K_2SO_4+Fe(CN)_6K_2Fe+6HCN.$$

Pure hydrocyanic acid is one of the most powerful and rapid of known poisons. When a small quantity of the vapor of the pure substance is drawn into the lungs instant death ensues, small quantities produce

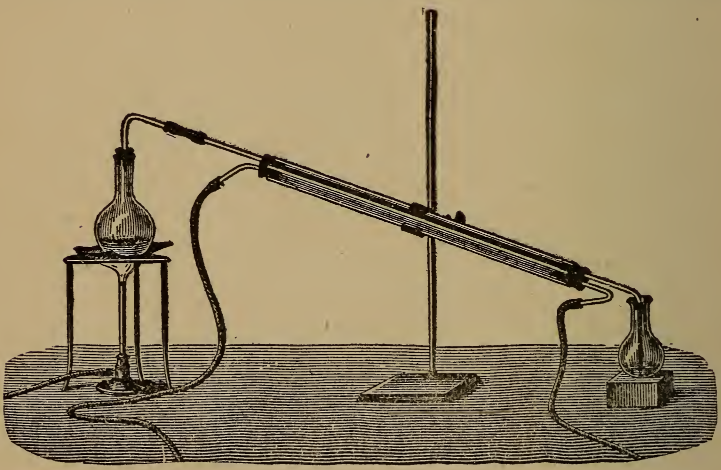

Fig. 22.

headache, giddiness, nausea, dyspnœa, and palpitation. A few drops brought into the eye of a dog kill it in thirty seconds, whilst an internal dose of 0.05 grain is usually sufficient to produce fatal effects upon the human subject, but cases have been known in which 0.1 grain has been taken without death ensuing. As antidotes, ammonia and chlorine water have been proposed, and these appear to be efficacious, although we are unable to explain their mode of action, for

ammonia under ordinary conditions only forms ammonium cyanide, and chlorine cyanogen chloride, both of which are bodies as poisonous as prussic acid itself. Hydrocyanic acid is used as a medicine, and is a constituent of several preparations, such as laurel water and bitter-almond water which are obtained by distilling the leaves of the common laurel, or bitter almonds with water.

Hydrogen, Boron and Oxygen Compounds.

Boric or Boracic Acid: Formula—H_3BO_3.

Boric acid is manufactured from minerals, such as borocalcite, which occurs in considerable quantities in the nitre beds of Peru and Chili. It is likewise prepared from the natural borax or tincal, which was first obtained from the basins of dried-up lagoons in Central Asia, and has been found in the borax lake in California in such quantities that the amount there obtained is sufficient to supply the whole demand of the United States. For the purpose of preparing boric acid from these sources, the minerals are dissolved in hot hydrochloric acid, the boric acid, which separates out on cooling, being recrystallized from hot water.

Boric acid crystallizes from aqueous solution in shining six-sided laminæ unctuous to the touch, and belonging to the triclinic system.

Boric acid is a weak acid, and its cold saturated solution colors blue litmus tincture a wine-red color like carbonic acid.

When boric acid is evaporated with an excess of concentrated phosphoric acid, and the dry residue treated with water in order to separate the phosphoric

acid, a white amorphous mass is left, which possesses the composition BPO_4. This substance is infusible, it is not attacked by strong acids, but dissolves in aqueous potash. The existence of these compounds points to the conclusion that boron trioxide possesses feebly basic properties, resembling, in this respect, alumina, Al_2O_3, which also sometimes acts as a weak acid, and sometimes as a weak base, forming a corresponding phosphate, $AlPO_4$.

Hydrogen, Bromine and Oxygen Compounds.

Hypobromus Acid, HBrO and Bromic Acid, $HBrO_3$.

Hypobromus Acid: Formula—HBrO.

This acid and the corresponding salts, termed hypobromites, are formed, in a similar manner to hypochlorous acid, by the action of bromine on certain metallic oxides. Thus if bromine water be shaken up with mercuric oxide, and if the yellow liquid thus formed be treated successively with bromine and the oxide, a solution is obtained which contains in every 100 c.c. 6.2 grams of bromine combined as hypobromous acid, the reaction being as follows,

$$HgO+2Br_2+H_2O=2HBrO+HgBr_2.$$

The greater part of the hypobromous acid contained in this strong solution is decomposed on distillation into bromine and oxygen. It can, however, be distilled in vacuo at a temperature of 40° without undergoing this change.

Aqueous hypobromous acid is a light straw-yellow colored liquid, closely resembling in its properties hy-

pochlorous acid, acting as a powerful oxidizing agent and bleaching organic coloring matters.

If bromine be dropped very slowly into a cooled solution of an alkali hydroxide, a hypobromite is formed along with bromide, but it is very unstable, and changes quickly into bromate.

By the action of bromine on lime, a substance similar to bleaching powder is formed, and this salt was formerly termed bromide of lime.

Bromic Acid: Formula—$HBrO_3$.

When bromine is dissolved in hot caustic potash or soda, a colorless solution is produced which contains a mixture of a bromide and a bromate, thus

$$3Br_2+6KHO=5KBr+KBrO_3+3H_2O.$$

The sparingly soluble potassium bromate may be easily separated from the very soluble bromide by crystallization. Potassium bromate is also formed when bromine vapor is passed into a solution of potassium carbonate which has been ensaturated with chlorine gas.

Solutions of alkali bromides are converted into bromates by electrolysis, the yield being almost quantitative if a little potassium chromate be added, in the case of the bromides of the alkaline earths some hypobromite is formed as well as bromate if the solution be kept cold.

Free bromic acid is formed when chlorine is passed into bromine water, thus

$$Br_2+5Cl_2+6H_2O=2HBrO_3+10HCl.$$

The acid is, however, best obtained by the decomposition of the slightly soluble silver bromate. This

salt is thrown down on the addition of silver nitrate to a solution of a soluble bromate, the precipitate thus prepared is well washed with water and then treated with bromine, bromic acid remains in solution and the insoluble silver bromide is thrown down,

$$5AgBrO_3+3Br_2+3H_2O=5AgBr+6HBrO_3.$$

Hydrochloric and hydriodic acids decompose bromic acid in a similar manner with formation of the chloride or iodide of bromine.

The bromates are as a rule sparingly soluble in water, and decompose on heating into oxygen and a bromide, but no per-bromate is formed in the process.

Hydrogen, Nitrogen and Oxygen Compounds.

Nitric Acid, HNO_3 and Nitrous Acid, HNO_2.

Nitric Acid or Hydrogen Nitrate: Formula—HNO_3.

Nitre, or potassium nitrate, is generally formed by the gradual oxidation of nitrogenous animal matter in presence of the alkali potash. Spring water, especially the surface well-water of towns, frequently contains nitrates in solution, owing to water passing through soil containing decomposing animal matters, which by oxidation yield nitrates. For this reason, water containing nitrates is unfit for drinking purposes. Potassium nitrate, KNO_3, commonly called saltpetre, occurs as an incrustation on the soil in various localities, especially in India, and sodium nitrate, $NaNO_3$, or Chili saltpetre, is found in large beds on the coast of Chili and Peru. Nitric acid is obtained by heating nitre, KNO_3, with sulphuric acid, or hydrogen sulphate, H_2SO_4, when nitric acid, HNO_3, and

hydrogen potassium sulphate, $HKSO_4$, are formed. The decompositions here effected may serve as a type of a very large number of chemical changes classed as double decompositions. These may all be represented as consisting in an exchange between two elements, or groups of elements, thus, in the case in question, one atom of the hydrogen in sulphuric acid changes place with one atom or its equivalent of potassium in the nitre. These double decompositions may be represented in the form of an equation, in which one side signifies the arrangement and relative weights of the elements before combination, the other the arrangement and relative weights of the same elements after the chemical change has taken place, thus

$$KNO_3+H_2SO_4=HNO_3+HKSO_4$$

or, Nitre and Sulphuric Acid give Nitric Acid and Hydrogen Potassium Sulphate. The relative weights of the elements and compounds entering into the decomposition are easily ascertained when we remember that each symbol expresses not merely the nature of the element, but also the relative weight with which it combines, and that the combining weight of a compound is the sum of the combining weights of its constituents. The numbers expressed by the above equation are

$$\begin{array}{cccccccccccc} K & N & O_3 + & H_2 & S & O_4 = & H & N & O_3 + & H & K & S \; O_4 \\ 39.1+ & 14+ & 48+ & 2+ & 32+ & 64= & 1+ & 14+ & 48+ & 1+ & 39.1+ & 32+64 \end{array}$$

$$101.1 \quad + \quad 98 \quad = \quad 63 \quad + \quad 136.1$$

We may express these double decompositions perhaps

more clearly if we represent the actual exchange of hydrogen for potassium, by a straight line, thus

$$\begin{array}{r|l} H & H\ SO_4 \\ NO_3 & K \end{array}$$

This signifies that, if we require 63 parts by weight of nitric acid, we shall require to take exactly 101.1 parts of nitre and 98 parts of sulphuric acid, and that we shall have 136.1 parts of hydrogen potassium sul-

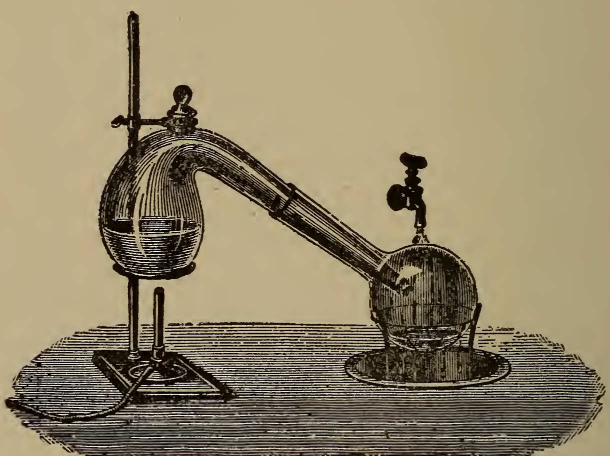

Fig. 23.

phate formed. Knowing these numbers, it is easy to calculate the proportions of ingredients needed to produce any given quantity of nitric acid.

Nitric acid is prepared on a small scale by placing about equal weights of nitre and sulphuric acid in a stoppered retort, which is gradually heated by a Bunsen burner, as in Fig. 23. The nitric acid formed distills over, and may be collected in a flask cooled with water. On a large scale this substance is prepared in

iron cylinders, into which the charges of nitre and acid are brought, the nitric acid being collected in large stoneware bottles.

Nitric acid thus obtained is represented by the formula HNO_3, it is a strongly fuming liquid, colorless when pure, but usually slightly yellow from the presence of lower oxides of nitrogen. In specific gravity it is 1.51 at 18°, it does not possess a constant boiling point, as it gradually undergoes decomposition by boiling and becomes weaker. If mixed with water, and distilled under the ordinary atmospheric pressure, the residual acid is found at last to attain a fixed composition, boiling constantly at 120.5°, containing 68 per cent. of HNO_3, and possessing a specific gravity of 1.414. When mixed with less water, a stronger acid than this comes over; when mixed with more water, a weaker one first distills over till this constant composition is attained. Nitric acid contains 76 per cent. of oxygen, with some of which it easily parts, hence it acts as a strong oxidizing agent. This is seen when we bring a small quantity of metallic copper or tin into this liquid diluted with a little water, red fumes are immediately given off, and the metals are oxidized, for the same reason nitric acid bleaches indigo solution, oxidizing, and therefore destroying, the coloring matter. This reaction, and the formation of red fumes in presence of metallic copper, etc., serve as modes of detecting the presence of nitric acid. One of the most delicate tests for this acid consists in adding to the liquid to be tested an equal volume of strong sulphuric acid, well cooling the mixture. and then carefully pouring on to its surface a solution of ferrous sulphate, $FeSO_4$. A black ring is produced where the two layers of liquid meet if any nitric acid

be present. Nitric acid forms, with metallic oxides, by the process of double decomposition, a numerous family of salts called nitrates. These are nearly all soluble in water, and many of them are largely used in the arts for various purposes.

In nitric acid we have the first example of a series of important compounds known as acids. Most of the acids are soluble in water, they possess an acid taste, and have the property of turning blue litmus-solution red. All acids contain hydrogen, combined either with an element, or with a group of elements which almost always contains oxygen. These acids may be regarded as water in which parts of the hydrogen is replaced by the oxygenated group of atoms. When the rest of the hydrogen of an acid is replaced by a metal, as for instance when sulphuric acid acts upon zinc, the acid character of the substance disappears, and a salt, called zinc sulphate, is formed. Salts are likewise produced when certain hydroxides and oxides are brought into contact with acids, thus if the solution of potassium hydroxide (caustic potash), obtained by the action of the metal potassium on water, is added to nitric acid, the alkaline or caustic properties of the hydroxide as well as the sour taste of the nitric acid disappear at a certain point, the solution becomes neutral, that is, it does not change the color of either blue or red litmus, and the salt potassium nitrate is contained in the liquid. The soluble hydroxides which thus act upon acids are termed alkalies, and have the power of turning red litmus-solution blue. In the same way many metallic oxides, called basic oxides or bases, act upon acids to form salts, thus silver oxide dissolves in nitric acid,

and neutralizes its acid character, forming soluble silver nitrate.

Nitrous Acid: Formula—HNO_2.

Nitrogen trioxide dissolves in ice-cold water giving rise to a beautiful blue liquid, which contains nitrous acid. Nitrous acid is not known in the pure state, it being a very unstable substance, which even in aqueous solution rapidly undergoes decomposition when warmed, giving rise to nitric acid and nitric oxide gas; thus

$$3HNO_2 = HNO_3 + 2NO + H_2O.$$

This reaction is, however, a reversible one, nitric oxide yielding with nitric acid and water a small quantity of nitrous acid, the condition of equilibrium being reached when about nine molecules of nitric acid are present to one of nitrous acid. The salts of this acid, or the nitrites, are, on the contrary, very stable bodies. They are not only formed by the action of the acid upon oxides, but also by the reduction of nitrates and by the oxidation of ammonia. Thus, for instance, potassium nitrate, KNO_2, is formed either by fusing saltpetre, or, more easily, by heating this salt with lead or copper, thus

$$2KNO_3 = 2KNO_2 + O_2.$$

It is also formed to some extent by the action of sunlight on a sterilized solution of potassium nitrate. Nitrites also occur in nature. Thus the atmosphere contains small quantities of ammonium nitrite, and traces of nitrites have been detected in the juices of certain plants. All the normal nitrites are soluble in water, and most of them soluble in alcohol. The silver salt is

the nitrite which is most sparingly soluble in cold water, crystallizing out in long glittering needle-shaped crystals when the hot aqueous solution is cooled.

When sodium nitrite is reduced by sodium amalgam, nitrogen and ammonia are the chief products, when the solution is hot, in the cold, nitrous oxide, hydroxylamine, and hyponitrous acid are formed, the relative amounts varying with the concentration.

The nitrites deflagrate when thrown on to glowing carbon, as do the nitrates. They can, however, be distinguished from the latter salts by the action of dilute acids, which produce an evolution of red fumes from the nitrites but not from the nitrates. In a similar way aqueous solutions of the neutral nitrites become of a light brown color when mixed with a solution of ferrous sulphate, and this color deepens to a dark brown on the addition of acetic acid. In order to detect the presence of a nitrite in dilute solution in the absence of other oxidizing agents, iodide of potassium, starch paste, and dilute sulphuric acid are added. The latter acid sets free the nitrous acid, and this instantly decomposes the iodide with liberation of iodine.

Nitrous acid is quantitatively oxidized by potassium per manganate and by hydrogen peroxide to form nitric acid, and is reduced to ammonia by boiling with ferrous sulphate and caustic soda. It may be estimated by any of these methods and also by its action on urea, which it oxidizes to nitrogen, carbon dioxide, and water. The free acid can, moreover, be titrated with caustic soda in dilute solution with methyl orange as indicator.

Hydrogen, Sulphur and Oxygen Compounds.

Sulphuric Acid, H_2SO_4 and Sulphurous Acid, H_2SO_3.

Sulphuric Acid or Oil of Vitriol: Formula—H_2SO_4.

This substance is the most important and useful acid known, as by its means nearly all the other acids are prepared, and also because it is very largely used in the arts and manufactures for a great variety of purposes. It has been truly said that the commercial prosperity of a country may be judged by the amount of sulphuric acid which it consumes.

Sulphuric acid was first prepared by distilling a compound of iron, oxygen, sulphur, and water, called ferrous sulphate or green vitriol. The acid thus obtained is known as fuming acid, and consists of a mixture of hydrogen sulphate and sulphur trioxide, $H_2SO_4+SO_3$. This plan of preparation has, however, long been superseded by the following more convenient method, which depends upon the fact that, although sulphur dioxide does not combine with free oxygen and water to form sulphuric acid, it is capable of taking up the oxygen when the latter is united with nitrogen in the form of nitrogen trioxide, N_2O_3, thus

$$SO_2+H_2O+N_2O_3=H_2SO_4+N_2O_2.$$

Sulphur dioxide, water and nitrogen trioxide yield sulphuric acid and nitric oxide.

The nitric oxide formed in this decomposition takes up another atom of oxygen from the air, becoming N_20_3, and this is again able to convert a second molecule of SO_2 with H_2O into H_2SO_4, being a second time reduced to N_2O_2, and ready again to take up another

atom of oxygen from the air. Hence it is clear that the N_2O_2 acts simply as a carrier of oxygen between the air and the SO_2, an indefinitely small quantity of this nitrogen trioxide being, therefore, theoretically able to convert an indefinitely large quantity of sulphur dioxide, water, and oxygen into sulphuric acid.

This process is carried on, on the large scale, in chambers made of sheets of lead (often a capacity of

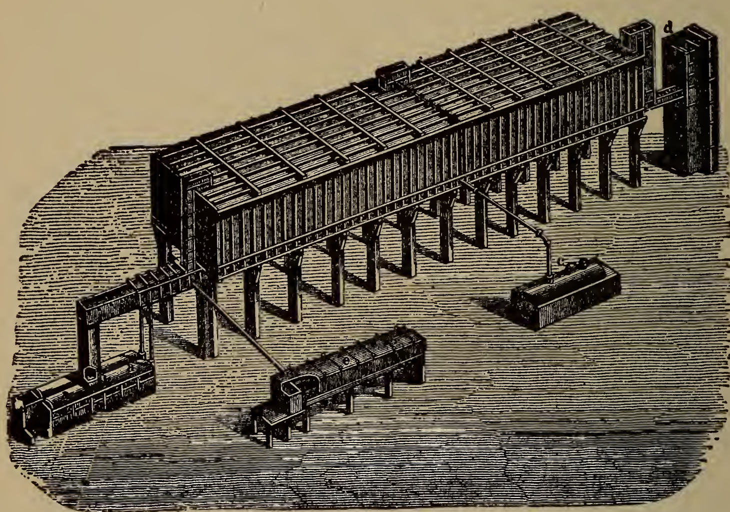

Fig. 24.

50,000 or 100,000 cubic feet) supported upon wooden beams and uprights, into which the above-mentioned materials are brought. Fig. 24 shows the arrangements for the manufacture of sulphuric acid. The leaden chambers, of which two are represented in the figure, are connected together by a wide leaden passage, and the gases passing from the first into the second chamber thus become thoroughly mixed. The sulphur diox-

ide is procured either by burning sulphur in a current of air, or by roasting a mineral called iron pyrites, a compound of sulphur and iron, FeS_2, in a suitable furnace **aa**. The sulphur of the pyrites burns away, and the gaseous product is led, together with atmospheric air, into the chamber, whilst ferric oxide, Fe_2O_3, remains behind in the furnace. A small stove **b** containing nitre, is placed in the central part of the furnace, where this salt is decomposed by the action of the sulphur dioxide, an alkaline sulphate being formed, whilst nitrous fumes pass with the other gases into the chamber. Jets of steam are also blown into the chamber at various points, from a boiler **c** and a thorough draft is maintained by connecting the end of the chamber with a high chimney not shown in the figure, but placed beyond the tower **d**. The fumes, gases, and air escaping from the chamber have to pass through the tower **d** and there meet with a jet of steam, by means of which the soluble acid vapors are nearly all condensed before reaching the chimney. The sulphuric acid, as it forms, falls on to the floor of the chamber, and, when the process is working properly, it is continually drawn off, attaining a specific gravity of about 1.60, the waste gases passing out of the chamber should contain nothing but nitrogen and small quantities of nitric oxide. In order to obtain from this weak chamber acid the pure sulphuric acid, H_2SO_4, the excess of water must be removed by evaporation: this is conducted, on the large scale, first, by heating the chamber acid in covered leaden pans **e**, until the specific gravity rises to 1.72, when the acid is known as the brown oil of vitriol of commerce, and then further concentrated in vessels of glass, or of platinum (as lead is attacked by the

strong acid), until its maximum strength and specific gravity is attained. The hydrogen sulphate thus obtained is a thick oily liquid boiling about 338°, and freezing at 10.8°, its specific gravity at 0° is 1.854. It combines with water with great force, absorbing moisture rapidly from the air. Hence it is used in the laboratory as a drying agent. Great heat is evolved when this acid is mixed with water, and care must be taken to bring these two liquids together gradually, otherwise an explosive combination may ensue. Many organic bodies, such as woody fibre and sugar, are completely decomposed and charred by strong sulphuric acid, whilst others, such as alcohol, oxalic and formic acid, are split up into other compounds by the withdrawal of the elements of water by this acid.

One molecule of hydrogen sulphate unites with one of water to form a compound, $H_2SO_4+H_2O$, which can be obtained pure by cooling a mixture of acid and water having a specific gravity of 1.78 down to 7° Centigrade, at which temperature rhombic crystals of the hydrated acid are formed. The sulphuric acid of commerce frequently contains large quantities of impurities, especially lead sulphate from the chamber, and frequently arsenic from the pyrites, and nitric acid, as well as the lower oxides of nitrogen. In order to free the acid from these impurities it must be distilled and subjected to other treatment, for a description of which the reader is referred to the larger treatises. At high temperatures sulphuric acid decomposes into sulphur dioxide, SO_2, oxygen, O, and water, H_2O. If a current of the acid be allowed to flow onto red-hot bricks, and the gases resulting from the decomposition passed through water, the sulphur dioxide will be com-

pletely absorbed, and a supply of pure oxygen obtained. Hydrogen sulphate is a dibasic acid, it contains two atoms of hydrogen, either or both of which can be replaced by an equivalent quantity of a metal. As with sulphurous acid, in the case of the alkaline metals, we have two salts, $KHSO_4$ and K_2SO_4. Barium and lead sulphates are insoluble in water, hence soluble salts of these metals are used as tests of the presence of a sulphate, a few drops of solution of barium chloride, for example, producing an immediate white precipitate of barium sulphate in water containing the merest trace of sulphuric acid or of a soluble sulphate. Calcium, strontium, and potassium-sulphates are but slightly soluble in water, whilst the other sulphates are easily soluble.

Some sulphates crystallize as anhydrous salts, such as K_2SO_4, potassium sulphate, $BaSO_4$, barium sulphate, and Ag_2SO_4, silver sulphate, whilst others require water to retain their crystalline form, and this water is termed water of crystallization. The crystals of iron sulphate or green vitriol, and of zinc sulphate or white vitriol, contain seven molecules of water in the solid form, while copper sulphate or blue vitriol requires but five molecules to preserve its crystalline form.

Sulphurous Acid: Formula—H_2SO_3.

This substance, like many other acids whose corresponding anhydrides are gaseous, is only known in aqueous solution. This solution smells and tastes like the gas and has a strongly acid reaction. Exposed to the light it is decomposed with formation of pentathionic acid. When an aqueous solution saturated at 3° is allowed to stand, crystals of a hydrate, H_2SO_3+

$6H_2O$, are obtained, and other hydrates of the formula H_2SO_3, $8H_2O$, H_2SO_3, $10H_2O$, and H_2SO_3, $14H_2O$ have been described. Sulphurous acid differs from the acids which have hitherto been described, inasmuch as it contains two atoms of hydrogen, both of which may be replaced by metals. It is therefore termed a dibasic acid, it forms two series of salts termed sulphites, in one of which only half of the hydrogen is replaced by a metal, and which may therefore be considered as being at once a salt and a monobasic acid, another in which the whole of the hydrogen of the acid has been replaced by a metal. The salts of the first series are termed acid sulphites and of the latter normal sulphites.

The acid sulphites of potassium and sodium are obtained by passing sulphur dioxide gas into caustic soda or caustic potash as long as it is absorbed. If, then, exactly the same quantity of alkali is added to this solution as was originally taken for the preparation, the normal salts are obtained. All the sulphites of the alkali metals are easily soluble in water, the normal sulphites of the other metals being either sparingly soluble or insoluble in water. They dissolve, however, in aqueous sulphurous acid and exist in such a solution as acid salts, but on evaporation they decompose with formation of the normal salt and sulphurous acid. The normal sulphites have no odor, and those which are soluble in water possess a sharp taste. They are readily detected by the fact that when they are mixed with dilute sulphuric acid they give off sulphur dioxide, and also that their neutral solutions give a precipitate with barium chloride which is soluble in dilute hydrochloric acid, whereas if nitric acid be added to this solution

and the mixture warmed, a precipitate of barium sulphate, formed by oxidation from the sulphite, is thrown down.

IRON.

Symbol—Fe. **Atomic Weight—55.6.**

Of all metals the most important to mankind, its uses were long unknown to the human race. The age of iron implements being preceded by those of bronze and stone. Pure metallic iron exists only in very small quantity on the earth's surface, almost entirely occurring in those peculiar structures known as meteoric stones, which possess an extra-terrestrial origin.

The process of obtaining iron from its ores is a somewhat difficult one, and requires an amount of knowledge and skill which the early races of men did not possess. The iron of commerce exists in three different forms, exhibiting very different properties, and possessing different chemical constitutions: **Wrought iron, Cast iron and Steel.**

The first is nearly pure iron, the second is a compound of iron with varying quantities of carbon and silicon, and the third a compound of iron with less carbon than that needed to form cast iron. The modes of manufacture of these three kinds of iron are essentially different, and will be best understood when the properties of the metal have been described.

Pure iron in the form of powder may be obtained by reducing the oxide, moderately heated in a current of hydrogen, it must, however, be retained in an atmosphere of hydrogen, as finely-divided iron takes fire and burns to oxide when exposed to the air. A button of pure iron may be prepared by exposing fine iron

wire mixed with some oxide of iron to a very high temperature in a covered crucible, the oxide retaining the traces of impurity which the wire contained. Iron has a bright white color, and is remarkably tough, though soft, an iron wire two mm. in thickness not breaking until weighted with 250 kilogs. The pure metal crystallizes in cubes. Iron which has been uniformly hammered exhibits, when broken, a granular and crystalline structure, this structure becomes, however, fibrous when the iron is rolled into bars, and the more or less perfect form of the fibre determines to a great extent the value of the metal. This fibrous texture of hammered bar iron undergoes a change when exposed to long-continued vibration, the iron returning to its original crystalline condition, and many accidents have occurred in the sudden snapping of axles, owing to this change from the fibrous to the granular texture. Wrought iron melts at a very high temperature, but as it becomes soft at a much lower point, it can be easily worked, as, when hot, it possesses the peculiar property of welding, that is, the power of uniting firmly when two clean surfaces of hot metal are hammered together.

Iron and certain of its compounds are strongly magnetic, but the metal loses this power when red hot, regaining it upon cooling. A solid mass of iron does not oxidize or tarnish in dry air, at the ordinary temperature, although iron powder takes fire spontaneously, but if heated it oxidizes, with the production of black scales of oxide, and when more strongly heated in the air, or plunged into oxygen gas, it burns, with the formation of the same black oxide. In pure water iron does not lose its brilliancy, but if a trace of car-

bonic acid is present, and access of air is permitted, the iron begins at once to oxidize at the surface, or to rust, forming a hydrated sesquioxide.

Iron and Oxygen Compounds.

Iron Monoxide, FeO, Iron Sesquioxide or Ferric Oxide, Fe_2A_3 and the Magnetic or Black Oxide of Iron, Fe_3O_4.

Iron Monoxide: Formula—FeO.

This substance has not been prepared in the pure state, owing to the readiness with which it absorbs oxygen, passing into the higher oxides. Hydrated ferrous oxide, FeH_2O_2, is thrown down as a white precipitate, when potash or soda is added to a soluble ferrous salt, this white precipitate can only be obtained in complete absence of oxygen, as it at once absorbs this gas, yielding a greenish-brown precipitate of a higher oxide. This oxide colors glass green, and gives the peculiar tint to common bottle-glass.

Iron Sesquioxide or Ferric Oxide: Formula—Fe_2O_3.

This oxide occurs native, as the minerals red hæmatite and specular iron ore, while combined with water, it forms brown hæmatite. It may be readily prepared artificially by heating ferrous sulphate to redness, or by adding solution of ammonia or caustic potash to a solution of a ferric salt, when the hydrated oxide falls down as a bulky brownish-red powder, which dissolves in acids, forming the ferric salts, when thus acted upon by sulphuric acid, ferric sulphate, Fe_23SO_4, is produced, and by hydrochloric acid, ferric chloride, Fe_2Cl_6.

Black Oxide of Iron: Formula—Fe_3O_4.

The magnetic oxide of iron occurs native, crystallized in octahedra, and, as the mineral loadstone, it constitutes one of the most valued ores of iron. It is the oxide formed when iron is oxidized at a high temperature in the air, in oxygen, or in aqueous vapor. A corresponding sulphide, Fe_3S_4, is also magnetic.

The Manufacture of Iron.

The oldest method of manufacturing wrought iron was to reduce it at once from the ore by heating in a furnace with charcoal or coal, and to hammer out the spongy mass of iron thus obtained. This plan can only be economically employed on a small scale and with the purest forms of iron ore, and has been superseded by a more complicated method, applicable, however, to all kinds of iron ore. This consists in the formation of cast iron as the first product, and the subsequent separation of the carbon and silicon which the cast iron contains. Cast iron is manufactured chiefly from clay ironstone, which generally occurs in masses, situated in the immediate neighborhood of a coal seam. The ironstone is first roasted, in which operation the carbonic acid is driven off, and ferric oxide formed, the ore afterwards being thrown, together with coal and limestone, into a blast furnace. It has the shape of a double cone, and is about fifty feet in height, and fifteen to eighteen feet in width at the broadest part. The furnace is closed at the bottom, the air necessary for the maintenance of the combustion being supplied in a powerful blast, blown through pipes called tuyeres. The mixture of fuel and ore, being put in at the top of

the furnace, is added continually as the burning mass sinks down and the molten mass is drawn off at the bottom, so that one furnace often does not stop working for several years. At the lowest part of the structure is the hearth, where the melted metal and fused slag collect, the former being occasionally tapped from the bottom of the hearth, and cast into pigs in moulds made in the sand, while the lighter slag, which swims on the surface of the metal, runs continually out from an opening at the upper part of the hearth.

The first chemical change which the roasted iron ore, or impure ferric oxide, undergoes in its passage from the top to the bottom of the furnace, is its reduction to a porous mass of metallic iron, by the carbonic oxide gas proceeding from the lower layers of burning coal. The temperature of this portion of the furnace is, however, much too low to melt the iron, and it therefore sinks down unchanged, together with the limestone, until it reaches a point at which the heat is greater. Here the second change occurs, the sand, and other impurities of the ore unite with the limestone to form a fusible silicate or slag, whilst the heated metal, coming in contact with carbon, unites at once with it to form cast iron, a fusible compound, which runs down to the bottom of the furnace. This, in passing through the hottest portion of the furnace, reduces the silica, with which it meets, to silicon, and, combined with this, it forms cast iron.

The properties and appearance of cast irons vary much with the quantity of carbon and silicon which they contain, for cast iron is not a definite chemical compound of these elements with iron. Carbon is found in cast iron, as scales of graphite, giving rise to mot-

tled cast iron, and in combination, forming white cast iron. Sometimes sulphur and phosphorus are also found in cast iron, but these must be considered as impurities. A great saving of fuel in the working of blast furnaces has been effected by employing the heat of combustion of the waste gases, which usually escape and burn at the top of the furnace, to raise the temperature of the blast of air supplying the furnace. The gases are collected at the top of the furnace by a hood and pass down an iron pipe, which is carried down to the furnaces in which the gases are burnt.

In order to obtain wrought from cast iron, the latter must undergo the processes of refining and puddling. These consist essentially in burning out the carbon, silicon, sulphur, and phosphorus, by exposing the heated metal to a current of air in a reverberatory furnace. The melted cast iron becomes first covered with a coat of oxide, and gradually thickens so as to allow of its being rolled into large lumps or balls. During this process the whole of the carbon escapes as carbonic oxide, and the silicon becomes oxidized to silica, which unites with the oxide of iron, and forms a fusible slag, any phosphorus or sulphur contained in the pig iron is also oxidized in this process. The ball is then hammered to give the metal coherence, and to squeeze out the liquid slag, and the mass is afterwards rolled into bars or plates.

Another interesting branch of the iron trade is the manufacture of steel. This substance is formed when bars of wrought iron are heated to redness for some time in contact with charcoal, the bar is then found to have become fine-grained instead of fibrous, the substance is more malleable and more easily fusible than

the original bar iron, and is found to contain carbon varying in amount from one to two per cent. Steel possesses several important properties, especially the power of becoming very hard and brittle when quickly cooled, which fits it for the preparation of cutting-tools, these are, however, generally made of bar-steel, which has been previously fused and cast into ingots.

A new and very rapid mode of preparing steel, which is of high industrial importance, is that known as the Bessemer process. This process consists in burning out all the carbon and silicon in cast iron by passing a blast of atmospheric air through the molten metal, and then in adding such a quantity of a pure cast iron to the wrought iron thus prepared as is necessary to give carbon enough to convert the whole mass into steel, the melted steel is then at once cast into ingots. In this way six tons of cast iron can at one operation be converted into steel in twenty minutes.

Iron, Carbon and Oxygen Compounds.

Carbonate of Iron: Formula—$FeCO_3$.

Ferrous Carbonate is an insoluble compound, and occurs largely as a mineral called spathose iron ore, which is isomorphous with calc-spar. It also occurs in a less pure form, constituting the clay ironstone, the ore of iron from which a large proportion of iron is prepared.

Iron and Chlorine Compounds.

Ferrous Chloride: Formula—$FeCl_2$.

When dry hydrochloric acid gas is passed over hot metallic iron, ferrous chloride and hydrogen are formed. The hydrated chloride is also produced when

iron is dissolved in aqueous hydrochloric acid, green crystals being deposited.

Iron, Sulphur and Oxygen Compounds.

Sulphate of Iron: Formula—$FeSO_4$.

Ferrous Sulphate, sometimes called green vitriol, is obtained by dissolving metallic iron, or ferrous sulphide, in sulphuric acid; and is also prepared by the slow oxidation of pyrites.

The solution thus obtained yields on evaporation large green crystals of the salt. It is largely used in the manufacture of several black dyes, and is one of the constituents of writing-ink. Like all the ferrous compounds, this salt easily takes up oxygen, producing a new salt called ferric sulphate.

Iron and Sulphur Compounds.

Ferrous Sulphide: Formula—FeS.

Ferrous Sulphide is an invaluable compound, formed by fusing equivalent quantities of sulphur and iron together, it is employed in the laboratory for the generation of sulphuretted hydrogen. A disulphide, called iron pyrites, is found in large quantities, and is much used in the production of sulphuric acid.

IODINE.

Symbol—I. **Atomic Weight—125.9.**

Iodine occurs combined with metals in sea-water, and is obtained from kelp, the ash of certain sea-weeds, in which it is found as the iodides of sodium and magnesium. Iodine is obtained from kelp by exactly the same process as that by which chlorine and bromine

are obtained from chlorides and bromides, by heating with sulphuric acid and manganese dioxide. Iodine is thus liberated in the form of a deep violet-colored vapor, which condenses to a dark grey solid, with bright metallic lustre. Iodine melts at 115°, and boils above 200°, and has a specific gravity of 4.95. It gives off a perceptible amount of vapor at the ordinary temperature, and possesses a faint chlorine-like smell. Water dissolves a very small quantity of iodine, but in presence of a soluble iodide it is freely dissolved, forming a deep red or brown solution. It is easily soluble in alcohol, giving a reddish-brown solution, and in carbon disulphide and chloroform, imparting to them a splendid violet color. Iodine does not possess such active properties as either of the preceding elements, its solution does not bleach organic coloring matters, and it is liberated from its compounds by both bromine and chlorine. Free iodine forms a remarkable compound with starch, of a splendid blue color, and by this means the minutest trace of this substance can be detected. To apply this test, one drop of potassium iodide solution is added to starch paste largely diluted with water, no blue color is observed until the iodine is set free by the addition of a drop or two of chlorine-water, when a deep blue coloration is instantly perceived. Iodine acts as a powerful poison, but, given in small quantities, it is much used as medicine.

The crude Chili saltpetre forms the chief source of iodine, to obtain the iodine the last mother-liquor obtained in the preparation of the sodium nitrate, which contains about 20 per cent of sodium iodate, is treated with a solution of sodium bisulphite, the iodine separates out in the solid form, and is filtered off and puri-

fied by resublimation, the vapors being condensed in a series of udells similar to those shown in Fig. 25.

In order to purify the commercial iodine it is washed with a small quantity of water, dried on porous plates, and resublimed. The only mode of obtaining chemically-pure iodine, free from every trace of chlorine and bromine, is to dissolve the commercial resublimed substance in iodide of potassium solution, and then to

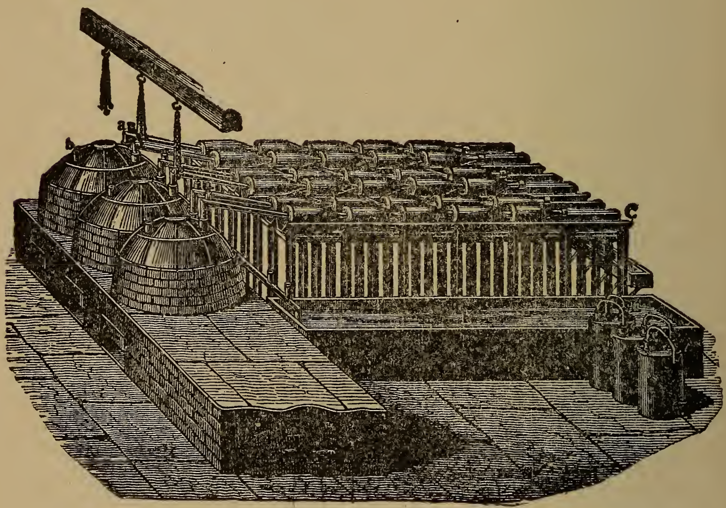

Fig. 25.

precipitate the iodine by water. The precipitate is well washed with water and then distilled with steam, the solid iodine in the distillate collected and dried in vacuo over solid nitrate of calcium, which is frequently changed, and distilled afterwards over solid caustic baryta to remove the last traces of water and of hydriodic acid.

Iodine is a bright, shining, crystalline, blackish-grey

solid, which is usually opaque, but may be obtained in transparent films by deposition on glass surfaces at a temperature of about —180°. The crystals when large possess almost a metallic lustre, it crystallizes by sublimation in the form of prisms or pyramids.

In its chemical properties iodine resembles chlorine and bromine, the two latter elements have the power of displacing iodine from its combination with metals, or electro-positive elements, thus:

$$2KI + Cl_2 = 2KCl + I_2.$$

The compounds of iodine with oxygen, or with electro-negative elements are, on the other hand, more stable than those of the other two elements. Thus iodine expels chlorine from the chlorates with formation of iodate and free chlorine,

$$2KClO_3 + I_2 = 2KIO_3 + Cl_2.$$

Iodine is very sparingly soluble in water, 1 part dissolving in 3,750 parts of water at 15° and in 2,200 at 30°. It dissolves readily in a solution of potassium iodide, forming a brown solution, in which, however, the iodine is probably not present as such, but as the unstable potassium triiodide. It also dissolves readily in a large number of other solvents, in some of which, such as chloroform, carbon bisulphide, and liquid hydro-carbons, it gives rise to violet solutions, whilst with others, such as alcohol, ether, organic acids and esters, and pyridine, it yields yellow or brown solutions. The cause of the difference of color of the various solutions has not yet been ascertained with certainty, but it seems most probable that in the violet solutions iodine is present as such, whilst the brown solutions contain an unstable compound of iodine with the solvent. The

first-named class of solutions when concentrated are black and opaque, in the case of carbon bisulphide, such a solution is diathermanous, allowing the invisible heating rays of low refrangibility to pass, but not the visible rays.

Although iodine, unlike chlorine and bromine, does not combine readily with hydrogen, it unites with many of the metals and non-metals with evolution of light and heat. Thus solid phosphorus, when brought into contact with iodine, first melts and then bursts into flame owing to the heat evolved in the act of combination, and powdered antimony takes fire when thrown into iodine vapor, antimony iodide being produced, while if the vapor of mercury be passed over heated iodine, immediate action occurs, the iodides of mercury being formed. When iodine is brought into contact with water and filings of iron or zinc, a violent reaction occurs, colorless solutions of the respective iodides resulting. The action of iodine upon the alkali-metals is analogous to that of chlorine and bromine. Sodium and iodine can be heated together without alteration, whilst if potassium be employed an explosive combination occurs.

LEAD.

Symbol—Pb. **Atomic Weight—204.85.**

Lead does not usually occur free in nature, all the lead of commerce is obtained from galena, or lead sulphide. The mode of reducing lead from this ore is a very simple one, the galena is roasted in a reverberatory furnace, with the addition of a small quantity of lime to form a fusible slag with any silicious mineral matter present in the ore. By the action of the air a

portion of the sulphide is oxidized to sulphate, whilst in another portion the sulphur burns off as sulphur dioxide, and lead oxide is left behind, after the lapse of a certain time the air is excluded and the heat of the furnace raised, the lead sulphate and oxide formed both decompose the remaining sulphide, giving off sulphur dioxide and leaving metallic lead behind.

Galena almost always contains small quantities of silver. Lead is a bluish-white colored metal, and so soft that it may be scratched with the nail, it may be drawn out to wire, or hammered into plate, but possesses little tenacity or elasticity, and a wire 2 mms. in diameter breaks with a load of 2 kilos. Lead melts at 334°, and at a higher temperature volatilizes, though not in quantity sufficient to enable it to be distilled.

The bright surface of the metal remains permanently in dry air, but it soon becomes tarnished in moist air, owing to the formation of a film of oxide, and this oxidation proceeds rapidly in presence of a small quantity of weak acid, such as carbonic or acetic. In pure water freed from air lead also preserves its lustre, but if air be present, lead-oxide is formed, and this dissolving slightly in the water a fresh portion of metal is exposed for oxidation. This solvent action of water upon lead is a matter of much importance, owing to the common use of lead water-pipes, and the peculiarly poisonous action of lead compounds upon the system when taken even in minute quantities for a length of time. The small quantity of certain salts contained in all spring and river waters exerts an important influence on the action of lead. The waters containing nitrates or chlorides are liable to contamination with lead, while those hard waters containing sulphates or

carbonates may generally be brought into contact with lead without danger, as a thin deposit of sulphate or carbonate is formed, which preserves the metal from further action. If the water contains much free carbonic acid, it should not be allowed to come into contact with lead, as the carbonate dissolves in water containing this substance. The presence of lead in water may easily be demonstrated by passing a current of sulphuretted hydrogen through a deep column of the acidified water, and noticing whether the liquid becomes tinged of a brown color, owing to the formation of lead sulphide.

Lead and Chlorine Compounds.

Lead Chloride: Formula—$PbCl_2$.

Lead Chloride is prepared by adding hydrochloric acid to a strong solution of lead nitrate, when a crystalline precipitate of lead chloride is formed. It dissolves in about thirty parts of boiling water, separating out in shining needles on cooling.

Lead and Chromium and Oxygen Compounds.

Lead Chromate: Formula—$PbCrO_4$.

Lead Chromate is a yellow insoluble salt, used as a pigment under the name of chrome-yellow.

Lead and Iodine Compounds.

Lead Iodide: Formula—PbI_2.

Lead Iodide is precipitated in the form of splendid yellow spangles, when hot solutions of potassium iodide and lead nitrate are mixed and allowed to cool.

Lead and Oxygen Compounds.

Lead Monoxide, PbO, Lead Dioxide, PbO_2 and Red Oxide or Red Lead.

Lead Monoxide or Litharge: Formula—PbO.

Lead Monoxide, or Litharge, is a straw-colored powder, obtained by heating lead in a current of air, it fuses at a red heat, forming scaly crystals termed Litharge or Massicot. Lead monoxide is soluble in caustic potash, and is deposited from a hot solution in the form of rhombic prisms. This oxide forms with acids the important series of lead salts, which are generally colorless, and of which the soluble ones act as violent poisons. Lead monoxide combines with silica to form an easily fusible silicate, or glass. Earthen crucibles in which the oxide is fused are rapidly attacked. A white hydrated oxide is obtained by precipitating a soluble salt of lead by caustic potash, and this if heated yields the oxide.

Lead Dioxide: Formula—PbO_2.

Lead Dioxide is a brown powder obtained by passing chlorine through the hydrated monoxide, or by digesting red lead with nitric acid. Lead dioxide does not form salts with acids. When heated it yields half its oxygen, and acted upon with warm hydrochloric acid, chlorine is evolved, and lead chloride is formed.

Lead Oxide:

Red Oxide, or Red Lead, a compound of the two last oxides, having the composition $2PbO+PbO_2$. It is obtained by exposing massicot to the air at a moderate

red heat, oxygen being absorbed. Red lead is chiefly used in glass-making and steam-fitting. When treated with dilute nitric acid the lead monoxide dissolves, forming soluble lead nitrate, leaving the puce-colored oxide behind.

Lead, Sulphur and Oxygen Compounds.

Lead Sulphate: Formula—$PbSO_4$.

Lead Sulphate is a white insoluble salt, which is found native, and is prepared artificially by adding sulphuric acid to a soluble lead salt.

LITHIUM.

Symbol—Li. **Atomic Weight—6.98.**

This metal is prepared by decomposing the fused chloride by electricity, it is of a white color, it fuses at 180°, and is the lightest metal known. The lithium salts were formerly supposed to be very rare, only being known to occur in three or four minerals, but spectrum analysis has shown that this is a widely-distributed substance. It occurs in small quantities in almost all waters, in milk, tobacco, and even in human blood. Lithium in its chemical relations stands between the class of alkaline and alkaline-earth metals, the hydrate, carbonate, and phosphate being only sparingly soluble in water. All the volatile lithium compounds impart a magnificent crimson tinge to the flame, and the spectrum of this flame exhibits the presence of one bright and very characteristic red line, by means of which the presence of the minutest trace of this substance can be detected with certainty and ease.

MAGNESIUM.

Symbol—Mg. **Atomic Weight—24.2.**

This metal occurs in large quantities as carbonate, along with calcium carbonate, in mountain limestone, and also in sea-water and certain mineral springs, as chloride and sulphate. The metal itself has only recently been prepared in quantity, it is best obtained by heating magnesium chloride with metallic sodium, sodium chloride and metallic magnesium being formed. This metal is of a silver-white color, and fuses at a low red heat, it is volatile, and may be easily distilled at a bright red heat, when soft it can be pressed into wire, and with care it may be cast like brass, although when strongly heated in the air it takes fire and burns with a dazzling white light, with the formation of its only oxide, magnesia. The light emitted by burning magnesium wire is distinguished for its richness in chemically active rays, and this substance is therefore employed as a substitute for sunlight in photography, and has been employed with success for photographing the interior of caverns.

Magnesium does not oxidize in dry air, it is only slowly acted upon by cold water, but more rapidly by hot water, it rapidly dissolves in sulphuric and hydrochloric acids, with evolution of hydrogen.

Magnesium, Carbon and Oxygen Compounds.

Magnesium Carbonate.: Formula—$MgCO_3$.

Magnesium Carbonate is an insoluble compound, occurring as a crystallized mineral termed magnesite. The white magnesia is a varying mixture of carbonate

and hydrate, made by precipitating a hot solution of magnesium sulphate with sodium carbonate. Magnesium sulphide is not formed in the wet way. Magnesium resembles in many respects the metals of the alkaline earths, but it may be distinguished from these by the solubility of the carbonate in ammonium chloride, as well as by the ready solubility of the sulphate in water. Magnesium forms an insoluble double phosphate with ammonia, and it is in this form the metal is usually estimated.

Magnesium and Chlorine Compounds.

Magnesium Chloride: Formula—$MgCl_2$.

Magnesium Chloride is a fusible salt obtained by evaporating magnesia dissolved in hydrochloric acid with an equal quantity of sal-ammoniac, on fusion, the latter salt volatilizes, and the magnesium chloride remains behind.

Magnesium and Oxygen Compounds.

Magnesium Oxide or Magnesia: Formula—MgO.

Magnesia is a light white amorphous infusible powder, obtained by heating the carbonate or nitrate, and is largely used in medicine, and known as calcined magnesia. It unites with acids to form the magnesium salts, but it does not possess strong alkaline reaction.

Magnesium, Sulphur and Oxygen Compounds.

Magnesium Sulphate: Formula—$MgSO_4+7H_2O$.

Magnesium Sulphate is a soluble substance known as Epsom Salts, it occurs in spring water, and contains seven atoms of water of crystallization, it is now large-

ly made from dolomite by separating the lime with the sulphuric acid. Magnesium sulphate forms, with the alkaline sulphates, double salts, in which the alkaline sulphate takes the place of one molecule of the water of crystallization.

MANGANESE.

Symbol—Mn. **Atomic Weight—54.6.**

Manganese occurs in nature as an oxide, and it can be obtained, though with difficulty, in the metallic state by heating the oxide very strongly with charcoal. The metal is of a reddish-white color, it is brittle, and hard enough to scratch glass. It decomposes water at the ordinary temperature, with evolution of hydrogen, it cannot be preserved in the air without undergoing oxidation, and must be kept under naphtha, or in a sealed tube, it is slightly magnetic, and, like iron, combines with carbon and silicon. Metallic manganese is not used in the arts, but an alloy of this metal and iron is now made on a large scale, and used in the manufacture of steel. Some of its oxides are used for the purpose of evolving chlorine from hydrochloric acid, and also for tinting glass a purple color.

Manganese and Oxygen Compounds.

Manganese Monoxide, MnO, Manganese Sesquioxide, Mn_2O_3 and Manganese Dioxide, MnO_2.

Manganese Monoxide: Formula—MnO.

Manganese Monoxide is a greenish powder, obtained by heating the carbonate in absence of air, it forms with acids a series of pink-colored salts, and rapidly

absorbs oxygen, passing into a higher state of oxidation. The hydrate is precipitated as a white gelatinous mass, when an alkali is added to a solution of manganous salt, this, however, rapidly becomes brown, owing to absorption of oxygen.

Manganese Sesquioxide: Formula—Mn_2O_3.

Manganese Sesquioxide exists in nature as braunite, and may be prepared artificially by exposing manganous oxide to a red heat. It forms a series of somewhat unstable salts, of which the manganese alum is one of the most interesting, being isomorphous with common alum, in which Mn_2O_3 is substituted for Al_2O_3.

Manganese Dioxide: Formula—MnO_2.

Manganese Dioxide is the common black ore of manganese, and is termed pyrolusite by mineralogists, it can be artificially formed by adding a solution of bleaching powder to a manganous salt. This substance yields one-third of its oxygen when heated to redness, forming the red oxide, and gives up half its oxygen when heated with sulphuric acid. It is largely used for the manufacture of chlorine.

Manganese, Potassium and Oxygen Compounds.
Manganic Acid: Formula—K_2MnO_4.
Permanganate of Potash: Formula—$KMnO_4$.

When an oxide of manganese is fused in the air with caustic alkali, a bright green mass is formed, which yields a dark green solution, this contains potassium manganate, or manganic acid, which may be crystallized, and is isomorphous with potassium sulphate and

chromate. If this green solution be allowed to stand, it slowly changes to a bright purple color, and hydrated manganese dioxide is deposited, hence its common name of mineral chameleon. It then contains a new salt in solution, a permanganate of potash, which may be obtained in the crystalline state by evaporation, and is isomorphous with potassium perchlorate. The presence of a few drops of acid at once effects this decomposition of the green solution.

The manganates and permanganates readily give up a part of their oxygen in presence of organic matter, and they are largely used as disinfectants.

MERCURY.

Symbol—Hg. **Atomic Weight—198.9.**

The principal ore of mercury is cinnabar, or vermillion, which occurs in California, and also in China and Japan. The metal is easily obtained by roasting the ore, when the sulphur burns off as the dioxide, and the metal volatilizes, and its vapor is condensed in earthen pipes. Mercury is the only metal liquid at the ordinary temperature, it freezes at —40°, in the solid state it is malleable. It boils at 350°, measured by the air thermometer, and gives off a slight amount of vapor at the ordinary temperature. Mercury when pure does not tarnish in moist or dry air, but when heated above 300° it slowly absorbs oxygen, passing into the red oxide, and it combines directly with chlorine, bromine, iodine, and sulphur. Hydrochloric acid does not attack mercury, sulphuric acid, on heating, forms sulphur dioxide and mercuric sulphate, nitric acid evolves nitric oxide, and forms mercuric nitrate. Mercury is largely

used in the process of extracting gold and silver from their ores and in the arts, for silvering mirrors, and other purposes. Mercury is deposited from its solutions upon metallic iron or copper, in the form of a grey powder, which becomes bright on burnishing. Mercury salts act as valuable medicines.

Mercury and Chlorine Compounds.

Mercuric Chloride: Formula—$HgCl_2$.

Mercuric Chloride, or Corrosive Sublimate, is prepared on a large scale by heating an intimate mixture of equal parts of mercuric sulphate and common salt. It is also formed when mercury burns in chlorine. It acts as a violent poison, it is soluble in water, fuses at 265°, and boils at 295°. When ammonia is added to a solution of mercuric chloride, the so-called white precipitate is thrown down. It is a chloride of mercury-ammonium.

Mercurous Chloride or Calomel: Formula—Hg_2Cl_2.

Mercurous Chloride or Calomel is generally prepared by heating a mixture of three parts of finely-divided metallic mercury with four parts of corrosive sublimate, the metal combines with half the chlorine of the corrosive sublimate, and one atom of calomel is formed. The calomel sublimes, and is deposited in a solid cake. It must be finely ground and well washed, in order to free it from any soluble mercuric chloride which may remain undecomposed. Calomel is a white powder, insoluble in water, it is decomposed by potash or ammonia. It is used largely in medicine.

Mercury and Oxygen Compounds.

Mercuric Oxide: Formula—HgO.

Mercury Monoxide, or Mercuric Oxide, is obtained by moderately heating the nitrate, or by heating the metal in the air for some time at a temperature of 300°. The oxide thus prepared appears as a dark-red crystalline powder, by precipitating it from a solution of the nitrate by caustic potash, it falls as an amorphous yellow powder.

Mercury and Sulphur Compounds.

Mercuric Sulphide or Cinnabar: Formula—HgS.

Cinnabar or vermillion occurs native, and may be prepared artificially by heating a mixture of sulphur and mercury. When precipitated from a solution of a mercuric salt by sulphuretted hydrogen, the sulphide falls as a black amorphous powder but on sublimation it becomes red and crystalline.

MOLYBDENUM.

Symbol—Mo. **Atomic Weight—95.3.**

The chief ore of this metal is molybdenum disulphide, a mineral in appearance resembling graphite. The metal is a grey substance, which oxidizes on heating in the air to molybdenum trioxide, a yellow powder which acts as an acid, forming with bases salts called molybdates. The compounds of molybdenum do not occur frequently, and are not used in the arts. Molybdic acid is, however, used as a reagent in the laboratory for detecting small traces of phosphoric acid.

NICKEL.

Symbol—Ni. Atomic Weight—58.4.

Nickel occurs in large quantities, combined with arsenic, as kupfernickel, also together with cobalt in speiss, and it is now prepared in considerable quantities for the manufacture of German silver, an alloy of nickel, zinc, and copper. Nickel is a white, malleable, and tenacious metal, it melts at a somewhat lower temperature than iron, and is strongly magnetic, but loses this property when heated to 350°.

Nickel and Oxygen Compounds.

There are two oxides of nickel, the monoxide, NiO, and the sesquioxide, Ni_2O_3. The former of these gives rise to the nickel salts, which possess a peculiar apple-green color. The monoxide is obtained by heating the nitrate or carbonate, or by precipitating a soluble nickel salt with caustic potash, and heating the apple-green hydrate, NiH_2O_2, which is thrown down. The sesquioxide is a black powder, prepared by adding a solution of bleaching-powder to a soluble nickel salt.

NITROGEN.

Symbol—N. Atomic Weight—13.9. Density—13.9.

Nitrogen is found in a free state in the atmosphere of which it forms about four-fifths by bulk, and occurs also in combination in many bodies such as ammonia, in the nitrates, and in many organic substances which form part of the bodies of vegetables and animals.

Pure nitrogen cannot readily be prepared from air,

as its separation from the gases of the helium group can only be effected by bringing the whole of the nitrogen into the combined state. It can, however, be obtained from nitrogen compounds in a variety of ways.

When a concentrated solution of ammonium nitrite is heated, the following reaction takes place:

$$NH_4NO_2 = N_2 + 2H_2O.$$

It is more convenient to employ a mixture of sodium nitrite with ammonium sulphate, which by double decomposition yields sodium sulphate and ammonium nitrite, the latter then decomposing into nitrogen and water. The addition of a little potassium bichromate prevents the formation of any nitric oxide, and the best results are obtained with a solution of 1 part of sodium nitrite, 1-2 part of ammonium sulphate, and 1 part of potassium bichromate, the gas being then washed through dilute sulphuric acid.

Pure nitrogen is also formed by the action of chlorine upon ammonia; thus

$$8NH_3 + 3Cl_2 = N_2 + 6NH_4Cl.$$

The chlorine evolved in a large flask passes into a three-necked Woulffe's bottle containing a strong aqueous solution of ammonia. The nitrogen gas which is here liberated is collected in the ordinary way over water, as shown in Fig. 26. Care must, however, be taken in this preparation that the ammonia is always present in excess, otherwise chloride of nitrogen may be formed, and this is a highly dangerous body, which explodes most violently.

Nitrogen is also evolved in many other reactions and decompositions, such as the action of sulphuric

acid on a mixture of ammonium nitrate and glycerol and the decomposition by heating of ammonium bichromate.

The oxygen of the air can readily be removed, yielding the mixture of gases formerly known as atmospheric nitrogen.

A small light porcelain basin is allowed to swim

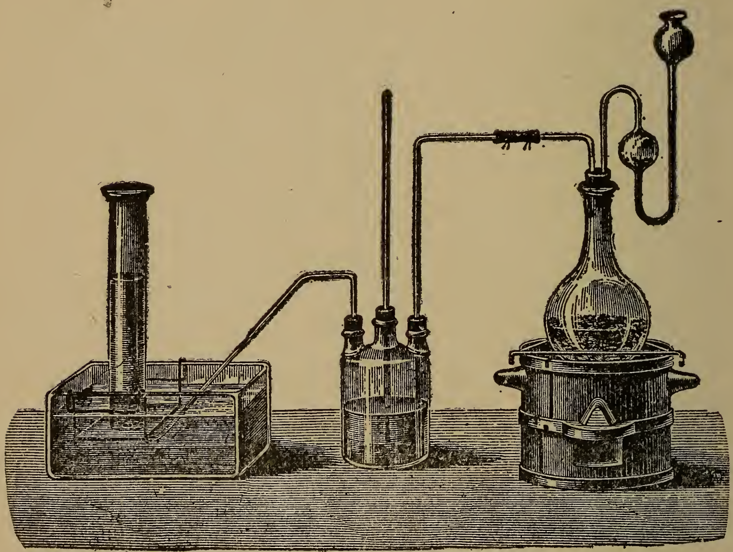

Fig. 26.

on the water of a pneumatic trough, a small piece of phosphorus brought into the basin and ignited, and the basin then covered by a large tubulated bell-jar as in Fig. 27. The phosphorus burns with the deposition of a white cloud of phosphorus pentoxide, which, however, soon dissolves, whilst on cooling, one-fifth of the contents of the bell-jar is found to be filled with water. The colorless residual gas is nitrogen, this

may be easily proved by first equalizing the level of the water inside and outside the bell-jar, after which, on opening the stopper and plunging a burning taper into the bell-jar, the flame is seen to be instantly extinguished. The nitrogen thus obtained is never perfectly pure, as, in addition to elements of the helium group, it always contains small quantities of oxygen which have not been removed by the combustion of the phosphorus. In presence of aqueous vapor, phos-

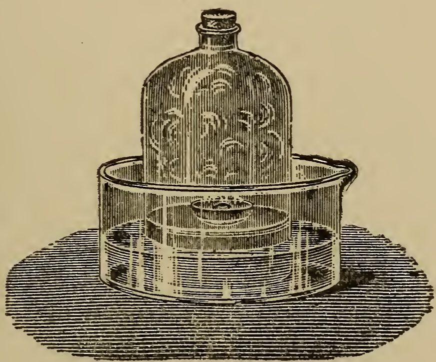

Fig. 27.

phorus slowly absorbed the oxygen of the air, at temperatures above 15°.

Pure nitrogen is a colorless, tasteless, inodorous gas, which is distinguished by its inactive properties, hence it is somewhat difficult to ascertain its presence in small quantities. As has been said, it does not support combustion, nor does it burn nor render lime-water turbid. It combines directly with but very few nonmetals, although indirectly it can easily be made to form compounds with most of these elements, and many of its compounds, such as nitric acid, ammonia

and chloride of nitrogen, possess characteristic and remarkable properties. It combines directly with a number of the metals, such as lithium, calcium, barium, and magnesium, yielding compounds termed nitrides.

Nitrogen forms an essential constituent of a very large number of animal and vegetable substances and is necessary for the maintenance of animal and vegetable life. It was however for a long time believed that members of the vegetable kingdom were unable to take up the free nitrogen of the air, and that they were dependent for their supply of this element on combined nitrogen contained in the atmosphere and in the soil chiefly in the form of nitric acid and ammonia. Certain leguminous plants, such as the white lupine, when grown in air free from ammonia and other nitrogen compounds, contain more nitrogen than was originally present in the seed and in the soil in which they were sown, and they must, therefore, have obtained the excess from the nitrogen of the air, since then it has been shown that certain algæ, fungi, and mosses behave in a similar manner, and it is not improbable that further investigation will show that this property is possessed by many other kinds of plants. This assimilation by plants is brought about by the action of certain micro-organisms, termed bacteroids, which are found in nodules on the roots, and absorb the free nitrogen from the air, forming compounds of nitrogen which are then assimilated by the plant. Other classes of micro-organisms also play a considerable part in the assimilation of nitrogen by plants, as many of the latter are incapable of directly assimilating ammonia, although they can take up nitric acid, the micro-organisms in the soil bring about the conversion

of ammonia into nitric acid, one species of these converting the former into nitrous acid, and another converting the nitrous into nitric acid.

Nitrogen and Chlorine Compounds.

Nitrogen Chloride: Formula—NCl_3.

Chloride of nitrogen is a thin yellowish oil which evaporates quickly on exposure to the air and possesses a peculiar smell, the vapor attacking the eyes and mucous membrane violently.

The following method is employed for preparing this dangerous substance in small quantities, and for showing its explosive properties without risk. A flask of about two litres capacity, having a long neck, is filled with chlorine, and placed mouth downwards in a large glass basin filled with warm saturated solution of sal-ammoniac as shown in Fig. 28. Below the neck of the flask is placed a small thick leaden saucer, in which the nitrogen chloride is collected. The solution of sal-ammoniac absorbs the chlorine, and, as soon as the flask is three parts filled by the liquid, oily drops are seen to collect on the surface inside the flask. These gradually increase, and at last drop one by one down into the leaden saucer. When a few drops have collected, the leaden saucer may be carefully removed by a pair of clean tongs, another being placed in its stead. A small quantity of the chloride of nitrogen may be exploded by touching it with a feather moistened with turpentine attached to the end of a long rod. Another drop of the oil may be absorbed by filtering paper, and when this is held in a flame a loud explosion likewise ensues.

The force of the explosion may be rendered still more evident by placing the flask in a strongly constructed box with glass sides, and when the drops of nitrogen chloride collect on the surface of the solution, passing in some turpentine from a stoppered funnel, the latter is kept outside the box and is connected with

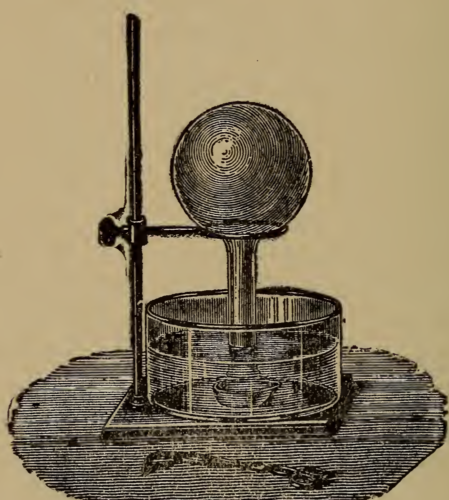

Fig. 28.

it by means of rubber and glass tube the end of which dips under the neck of the flask. When the turpentine reaches the surface of the solution in the flask a bright flash is seen, and a violent thunderlike explosion occurs, completely shattering the flask.

Nitrogen and Hydrogen Compounds.

Ammonia, NH_3—Hydrazine, N_2H_4 and Azoimide, N_3H.

Ammonia: Formula—NH_3. Molecular Weight—16.93. Density—8.45.

Ammonia is chiefly obtained from the decomposition of animal or vegetable matters containing nitrogen and hydrogen, being formed either gradually at the ordinary temperature, or quickly under the influence of heat. When horns, or clippings of hides, or coal is heated, ammonia is given off; hence ammonia was known as spirits of hartshorn. The name ammonia is derived from the fact that a compound containing ammonia, called sal-ammoniac, was first prepared by the Arabs in the deserts near the temple of Jupiter Ammon, by heating camels' dung. Guano, the dried excrement of sea-birds, and the urine of animals, likewise contain large quantities of ammonia. Ammonia and its compounds are now, however, mainly obtained from the ammoniacal liquors of gasworks. Coal contains about 2 per cent of nitrogen, which, when the coal is heated in close vessels, mostly comes off in combination with the hydrogen of the coal as ammonia. Hydrochloric acid is added to this ammoniacal liquor, and the solution evaporated, when the sal-ammoniac of commerce is obtained.

Ammonia may also be formed by the action of nascent hydrogen on dilute nitric acid, and when this acid is placed in contact with metallic zinc or iron, ammonia is formed.

Ammonia gas is best prepared by heating in a glass flask one part of sal-ammoniac, or ammonia hydrochlorate, and two parts of powdered quicklime.

Quicklime and sal-ammoniac give calcium chloride, and ammonia and water.

Ammoniacal gas is colorless, and possesses a most pungent and peculiar smell, by means of which it can be readily recognized, it is lighter than air, and it may be collected by displacement, the neck of the bottle intended to receive the gas being turned downwards, as in Fig. 29. A cylinder filled with quicklime is here

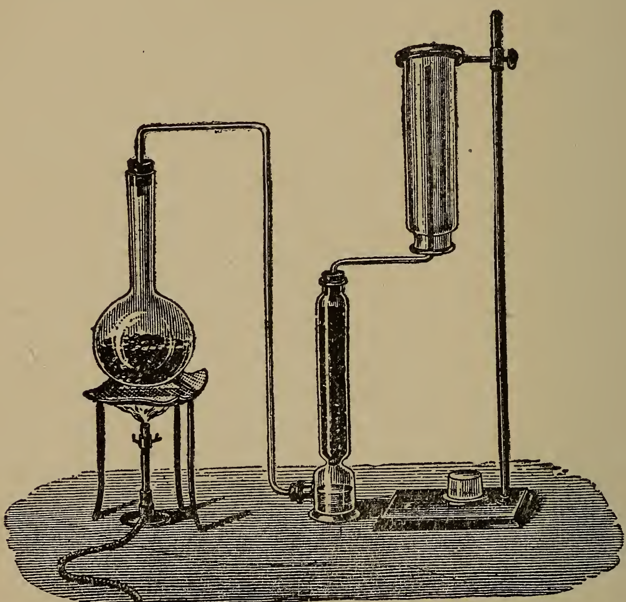

Fig. 29.

placed between the flask and the bottle, for the purpose of completely drying the ammonia. Ammonia may also be collected over mercury, but not over water, as it is extremely soluble in this liquid, one gram of water at 0° absorbing 0.877 gram, or 1149 times its volume, of ammonia, under a pressure of 760 mm., whilst at 20° the same weight of water absorbs 0.520 gram, or

681.1 times its volume, under the same pressure. The solution of ammonia gas in water is the common liquor ammoniæ of commerce. Ammonia gas, as well as the aqueous solution, possesses a strong alkaline reaction, turning red vegetable colors blue. It unites with the most powerful acids, forming compounds called the salts of ammonia, which closely resemble the salts of the alkaline metals, hence the name of the volatile alkali has been given to ammonia.

On exposure to a pressure of seven atmospheres at the ordinary temperature of the air, about 15° Centigrade, ammonia condenses to a colorless liquid, boiling at —38.5°, and this liquid, if cooled below —75°, freezes to a transparent solid. An elegant application of the principle of the latent heat of vapors has recently been made in the case of ammonia in M. Carre's freezing machine. This consists essentially of two strong iron vessels connected in a perfectly air-tight manner by a bent pipe, one of these vessels contains an aqueous solution of ammonia saturated with the gas at 0°. When it is desired to procure ice, the vessel containing the ammonia solution is gradually heated over a large gas burner, the other vessel being placed in a bucket of cold water. In consequence of the increase of temperature, the gas cannot remain dissolved in the water, and passes into the receiver, where, as soon as the pressure amounts to about 10 atmospheres, it condenses in the liquid form. When the greater part of the gas has thus been driven out of the water, the apparatus is reversed, the retort being cooled in a current of cold water, whilst the liquid it is desired to freeze is placed in the interior of the receiver. A reabsorption of the ammonia by the water now takes

place, and a consequent evaporation of the liquefied ammonia in the receiver. This evaporation is accompanied by an absorption of heat which becomes latent in the gas, hence the receiver is soon cooled far below the freezing point, and ice is produced around it.

The composition of ammonia may be ascertained by leading the gas through a red-hot tube, or passing a series of electric sparks through the gas, when it will be decomposed into nitrogen and hydrogen, which will be found to occupy together a volume twice as large as the ammonia taken, and to be mixed together in the proportions of three volumes of hydrogen to one volume of nitrogen.

Hydrazine or Diamide: Formula—N_2H_4.

Free hydrazine is difficult to prepare on account of the great stability of the hydrate, which is formed whenever the base is liberated in presence of water.

It is, however, best prepared by gradually adding the hydrate to anhydrous barium oxide, heating for some time at 110—120° and then distilling under 150—100 mm. pressure, the air of the apparatus being displaced by hydrogen so as to prevent the oxidation of the base. The distillate is not quite free from water and must therefore be again distilled over baryta. The free base is a colorless liquid, which boils at 113.5° and readily solidifies forming crystals which melt at 1.4°. The base is extraordinarily hygroscopic and mixes readily with water or alcohol, but is only very sparingly soluble in organic solvents generally. It fumes in the air and is readily inflammable, but not explosive. It acts as a most violent reducing agent, and bursts into flame when brought into contact with

chlorine, free nitrogen and hydrochloric acid being formed, on bromine and iodine it also reacts vigorously, forming the corresponding acids. It slowly oxidizes in the air and when strongly heated decomposes, the final products being ammonia and nitrogen,

$$3N_2H_4=N_2+4NH_3.$$

Free hydrazine dissolves many salts and yields solutions which conduct electricity, so that its action in this respect is comparable with that of water and ammonia.

Azoimide or Hydrazoic Acid: Formula—N_3H.

A convenient source of azoimide is the crude amidoguanidine obtained in the preparation of hydrazine, this substance is converted by nitrous acid into diazoguanidine nitrate, which on boiling with alkalies yields azoimide and cyanamide, the former may be isolated by acidification and distillation.

A further very interesting synthesis of azoimide from purely inorganic sources is by obtaining its sodium salt by the action of nitrous oxide on sodamide.

The water formed acts on a further molecule of sodamide yielding caustic soda and ammonia. The sodamide is obtained by passing ammonia over metallic sodium at a temperature of 150—250°, and as soon as all metallic sodium has disappeared, the steam of ammonia is replaced by one of nitrous oxide and continued till ammonia is no longer evolved, the product is then dissolved in water and distilled with dilute sulphuric acid.

Several other interesting modes of formation have also been discovered. A yield of 36 per cent. of the

theoretical amount is obtained when a benzene solution of nitrogen chloride is shaken with a solution of hydrazine sulphate, and caustic soda added at intervals. It is also formed when a mixture of molecular proportions of hydrazine and hydroxylamine dissolved in dilute sulphuric acid is oxidized by chromic acid or hydrogen dioxide. The silver salt may also be conveniently prepared by cautiously warming hydrazine sulphate with nitric acid and passing the gas evolved into silver nitrate solution.

Nitrogen and Oxygen.

The Air.

That the oxygen and nitrogen of the air are mechanically mixed and not chemically combined is seen from the following facts:

The quantities of nitrogen and oxygen in the air do not present any simple relation to the atomic weights of these elements, and the proportions in which they are mixed are variable.

On mixing oxygen and nitrogen gases mechanically in the proportion in which they occur in air, no contraction or evolution of heat is observed, and the mixture behaves in every way like air.

When air is dissolved in water, the proportion between the oxygen and nitrogen in the dissolved air is quite different from that of the undissolved air, the difference being in strict accordance with the laws of gas-absorption on the assumption that the air is a mixture. When water is saturated with air at any temperature below 30°, the following is the proportion of oxygen and nitrogen contained in the dissolved and the original air:

	Air dissolved in water.	Air undissolved in water.
Oxygen	35.1	20.96
Nitrogen	64.9	79.04
	100.0	100.00

If the air were a chemical combination of oxygen and nitrogen, such a separation by solution would be impossible.

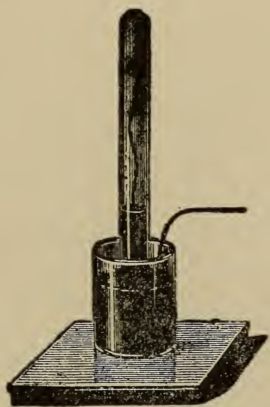

Fig. 30.

When liquefied air is allowed to boil, the nitrogen passes off much more rapidly than the oxygen, which could not take place if the two gases were chemically combined.

In order to show the composition of the air, the apparatus Fig. 30 is used. This consists of a calibrated and divided glass tube filled to a given point with air over mercury. Into this is introduced a small piece of phosphorus supported upon a copper wire. Gradually all the oxygen is absorbed and the mercury rises in the

tube. After a while the volume of residual gas is read off, and, corrections having been made for temperature and pressure, it is found that 100 volumes of the air contain about 21 volumes of oxygen.

Another less exact but more rapid method of exhibiting the same fact is carried out by help of the

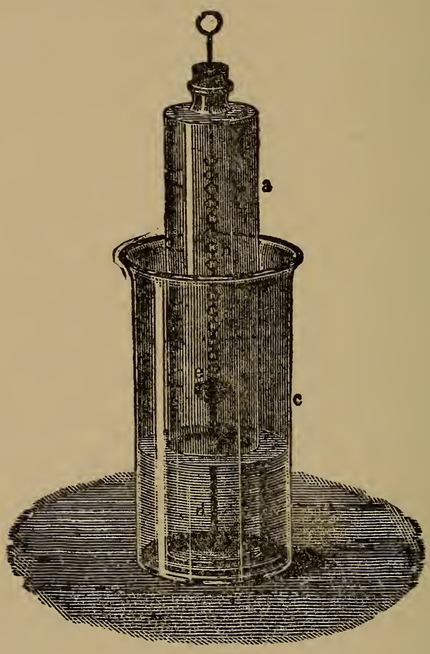

Fig. 31.

arrangement shown in Fig. 31. In the beaker-glass **c** is placed the iron stand **d** carrying the iron cup **e**, containing a small piece of phosphorus. Over this stand is placed the cylinder **a**. The upper part of this cylinder is graduated into five equal divisions, and water is poured into the beaker-glass until the level reaches the first division. The phosphorus in the cup is ignited

by dropping down on it a chain which has been heated in a flame. The phosphorus then burns, the fumes of phosphorus pentoxide are absorbed by the water, and, when the gas has cooled, and the pressure been equalized by bringing the level of the water outside up to that inside the cylinder, four-fifths of the original volume of the air remain unabsorbed.

Analyses of air collected in various parts of the globe, made with the greatest care have shown that the relative quantities of oxygen and nitrogen remain the same, or very nearly the same, from whatever region the air may have been taken. So that whether the air be derived from the tropics or the arctic seas, from the bottom of the deepest mine or from an elevation of 20,000 feet above the earth's surface, it contains from 20.9 to 21 volumes of oxygen per cent.

When we know the composition of air by volume, and the relative densities of the two constituent gases, 14 for nitrogen and 16 (actually 15.88, when hydrogen is 1), for oxygen, we can calculate its composition by weight, we thus find that in 100 grams of air, 23.16 grams of oxygen are mixed with 76.84 grams of nitrogen. It is important to control this calculation by experiment, for this purpose a large glass globe furnished with a stopcock is rendered vacuous by the air-pump and then weighed, a tube of hard glass filled with copper turnings and also furnished with stopcocks is likewise weighed. This tube is then heated to redness in a long tube-furnace, and connected at one end with the empty flask, at the other with a series of tubes filled with caustic potash and sulphuric acid, for the purpose of completely freeing the air passing through them from carbonic acid and aqueous vapor, the cocks

are then slightly opened, and air allowed to pass slowly through the purifiers into the hot tube, where it is completely deprived of oxygen by the hot metallic copper, which is thereby oxidized, the nitrogen passing on alone into the empty flask. After the experiment is concluded, the cooled tube is again weighed, and the increase over the former weighing gives the quantity of oxygen, whilst the increase in weight of the globe gives the nitrogen. The mean of a large number of experiments thus made shows that 100 parts by weight of air contained 23 parts by weight of oxygen and 77 of nitrogen.

In addition to the two above-mentioned gases, the air contains several other important constituents, especially carbonic acid gas, aqueous vapor, and ammonia gas. The carbonic acid gas of the air plays an important part in the phenomena of vegetation, this gas being the source from which plants obtain the carbon they need to form their tissues. The quantity of carbonic acid present in the air is very small compared with the quantities of oxygen and nitrogen, being only about 4 volumes to 10,000 of air, nevertheless the absolute quantity of this gas contained in the whole atmosphere is enormously large, about 3,000 billion kilos. The quantity of carbonic acid contained in the air can be found by drawing a known volume of perfectly dry air, not less than 20 litres, through weighed tubes containing caustic potash, the increase in weight of the tubes gives the weight of carbonic acid contained in the air drawn through. Hail is caused by the congelation of raindrops in passing through a stratum of air below the freezing point. The quantity of rain thus deposited is very large: 1 cubic metre of air saturated

with moisture at 25° Centigrade contains 22.5 grams of water, and if the temperature of this air be reduced to 0° Centigrade, it will then be capable of retaining only 5.4 grams of water vapor, hence 17.1 grams of water will be deposited as rain. Instruments for ascertaining the degree of moisture or humidity of the air are termed hygrometers.

The deposition of dew is caused by the rapid cooling of the earth's surface by radiation after sunset, and by the consequent cooling of the air near the ground below the temperature at which it begins to deposit moisture.

The amount of aqueous vapor contained in the air at any time can be determined by the apparatus used for the estimation of the carbonic acid, for the moisture must be removed from the air before the carbonic acid can be absorbed, and the increase in weight of the tubes filled with pumice-stone moistened with sulphuric acid gives the weight of aqueous vapor. In general the air contains from 50 to 70 per cent of the quantity necessary to saturate it. If the quantity be not within these limits, the air is either unpleasantly dry or moist.

The next important constituent of the air is ammonia, which is a compound of nitrogen and hydrogen, and only exists in comparatively very minute quantities, about 1 part in 1,000,000 of air. Nevertheless it plays a very important part, as it is mainly from this ammonia that vegetables obtain the nitrogen which they need to form their seeds and fruit, for it appears that all plants have not the power of assimilating the free nitrogen of the atmosphere. Other substances which occur in the atmosphere in very small quantities

may be considered as accidental impurities. Amongst them, volatile organic matter is the most important, as probably influencing to a great extent the healthiness of the special situation. We become aware of the existence of such organic putrescent substances when entering a crowded room from the fresh air, and it is probable that the well-known unhealthiness of marshy and other districts is owing to the presence of some organic impurity. At present, however, we possess but little certain knowledge on this subject. Ozone is also present in fresh air, but generally absent in the close air of towns and dwelling-rooms, owing to its decomposition by the organic matter in such air, we do not know how it is formed in nature, unless it be by the discharge of atmospheric electricity.

Nitrogen and Oxygen Compounds.

Nitrogen Monoxide or Nitrous Oxide—N_2O,
Nitrogen Dioxide or Nitric Oxide—NO,
Nitrogen Trioxide—N_2O_3,
Nitrogen Peroxide—NO_2,
Nitric Anhydride—N_2O_5.

Nitrous Oxide: Formula—N_20. Molecular Weight—43.7. Density—21.85.

Nitrous Oxide is called laughing gas, because when it is mixed with air and inhaled, it produces a peculiar and transient intoxicating effect, and when inhaled in its pure state for a few minutes, it acts as an anæsthetic, and renders the person insensible to pain for a short time.

Nitrate of ammonia which is prepared by neutral-

izing nitric acid with ammonia, gives off nitrous oxide when decomposed by heat as in Fig. 32, thus

$$HNO_3+NH_3=(NH_4)NO_3$$
$$\text{and } (NH_4)NO_3=N_2O+2H_2O.$$

That is, ammonium nitrate yields nitrous oxide and water.

Nitrous oxide is a colorless inodorous gas possessing a slightly sweet taste, it is somewhat soluble in cold water, one volume of water at 0° dissolving 1.305

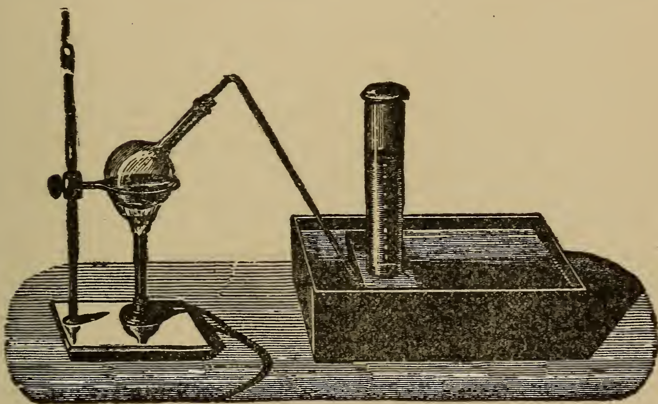

Fig. 32.

volumes of the gas, whilst one volume of water at 24° dissolves only 0.608 volume. Nitrogen monoxide differs from all such gases which we have previously considered, inasmuch as it liquefies when exposed either to great pressure or to an intense degree of cold. Thus, if it be brought under a pressure of about 30 atmospheres at 0°, or if it be cooled down to —88° under the ordinary pressure, it forms a colorless liquid, in other words, the pressure of nitrous oxide vapor or gas is 1 atmosphere at —88°, and 30 atmospheres at 0° Centi-

grade. If this liquid be cooled below —115°, it solidifies to a transparent mass. By the rapid evaporation of this liquid in vacuo, the lowest artificial temperature hitherto known has been attained, about —140° Centigrade.

A glowing chip of wood when plunged into nitrous oxide rekindles, and the wood continues to burn with a brighter flame than in the air, whilst phosphorus on burning in this gas evolves nearly as much light as in pure oxygen, a feeble flame of sulphur is, however, extinguished on bringing it into this gas, but if burning strongly it also continues to burn brightly. This is owing to the fact that the gas has to be decomposed into nitrogen, one volume, and oxygen, half a volume, before bodies can burn in it, and to effect this decomposition a tolerably high temperature is necessary.

Nitric Oxide: Formula—NO. Molecular Weight—29.8. Density—14.9.

A colorless gas obtained by acting upon copper turnings with nitric acid. Copper and nitric acid give copper nitrate, nitrogen dioxide, and water.

This substance has not been condensed to a liquid, in contact with oxygen it combines directly with this gas, forming red fumes which are readily soluble in water, and by this property it may be distinguished from all other gases. Although nitric oxide contains half its volume of oxygen, and more oxygen in proportion by weight than nitrous oxide, it does not easily support combustion, as it requires a high temperature for its decomposition, thus, ignited phosphorus, unless burning very brightly, is extinguished on plunging it into nitric oxide gas.

The composition of this gas may be determined as follows: One volume of nitrogen dioxide yields half a volume of nitrogen, as the weight of one volume of nitrogen dioxide is 15, the weight of oxygen contained in one volume of this gas is 15—7=8, or two volumes of nitrogen dioxide weigh 30, and are composed of one volume of nitrogen weighing 14, and one of oxygen weighing 16 (actually 15.88 when hydrogen is 1). Hence, in accordance with the law respecting the densities of compound gases, the formula of this oxide should be NO and not N_2O_2, the physical properties of the gas likewise, compared with those of nitrous oxide, seem to indicate that this latter has a more complicated constitution.

If N_2O_2 is used for the formula then this gas will be an exception to the law that the density of any compound gas is half its molecular weight. The density being 15, the molecular weight, according to this law, is 30 and the formula is NO. It is as though the single molecule N_2O_2 had split up into two molecules of NO. The name nitrogen dioxide is given to this gas because for the same weight of nitrogen it contains twice as much oxygen as nitrogen monoxide.

Nitrogen Trioxide: Formula—N_2O_3. Molecular Weight—75.5. Density—37.7.

This substance is prepared by mixing four volumes of dry nitrogen dioxide with one volume of oxygen, and cooling the mixture to —18°, the two gases combine to form red fumes, which condense to a volatile indigo-blue colored liquid, the same blue body is obtained by adding water to nitrogen peroxide and drying the distillate over calcium chloride. It is also

formed by the action of moderately strong nitric acid upon arsenic trioxide, with formation of arsenic acid. Arsenic trioxide and nitric acid and water yield nitrogen trioxide and arsenic acid.

Nitrogen trioxide dissolves in ice-cold water, forming a blue liquid, and containing nitrous acid or hydrogen nitrite, HNO_2, in solution, this compound is very unstable, and decomposes when the water is warmed, into nitric acid and nitric oxide.

The salts formed by nitrous acid are, however, not liable to such easy decomposition, potassium nitrite, KNO_2, is obtained by heating potassium nitrate, KNO_3, which loses one atom of oxygen, the same salt is produced when nitrogen trioxide is led into a solution of caustic potash.

Hence, nitrogen trioxide stands to the nitrites in the same position as nitrogen pentoxide to the nitrates. It will be noticed that nitric acid forms salts called nitrates, whilst nitrous acid gives rise to nitrites, this is an example of a general rule adopted in chemical nomenclature that if the specific name of an acid or hydrogen salt end in **ous**, the names of the corresponding metallic salts end in **ite**, whilst acids whose names end in **ic** form salts ending in **ate.**

Nitrogen Peroxide: Formula—NO_2. Molecular Weight—45.7. Density—22.8.

The red fumes which are formed when nitric oxide comes into contact with oxygen or air, consist chiefly of nitrogen peroxide. If one volume of dry oxygen be mixed with two volumes of dry nitric oxide and the red fumes which are produced led into a tube sur-

rounded by a freezing mixture, the peroxide condenses in the tube either as a liquid or in the form of crystals.

Nitrogen peroxide is also formed by the decomposition which many nitrates undergo when heated, and is usually prepared by strongly heating lead nitrate in a retort of hard glass as shown in Fig. 33.

This mode of preparation is, however, not very convenient, and a considerable loss of material occurs, as the oxygen which is evolved carries away some quan-

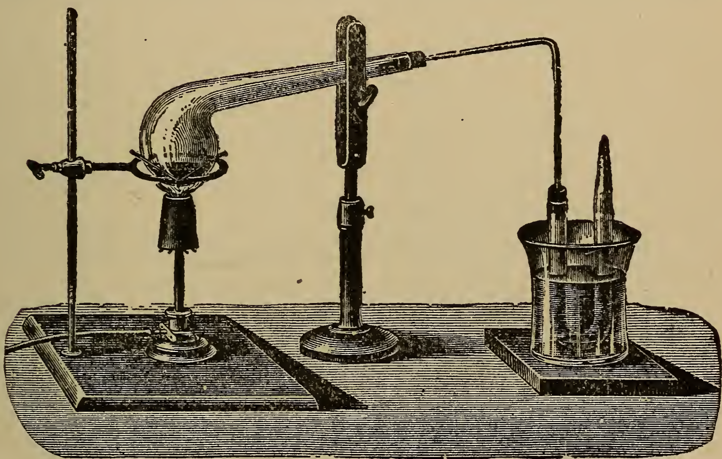

Fig. 33.

tity of the peroxide even when the tube into which the fumes are led is plunged into a freezing mixture.

The following method is free from the above objections. Arsenious oxide (white arsenic) in the form of small lumps is placed in a flask and covered with ordinary nitric acid, or, with a mixture of nitric acid of specific gravity 1.5 and half its weight of sulphuric acid, the red fumes, which are given off in quantity on gently heating, are led into a receiver surrounded by

a freezing mixture, where a mixture of trioxide and tetroxide of nitrogen collects. The mixture is freed from the trioxide by the addition of strong nitric acid and a large quantity of phosphorus pentoxide, and the tetroxide is then poured off from the syrupy layer, and distilled. The trioxide may also be converted into peroxide by a current of air.

Nitrogen peroxide may also be prepared by mixing the fumes obtained by the action of nitric acid of specific gravity 1.4 on arsenious oxide with oxygen, drying over calcium nitrate, condensing and redistilling.

Nitric Anhydride: Formula—N_2O5.

This compound is a white crystalline solid obtained by removing the elements of water from nitric acid by means of phosphorus pentoxide, a substance which has a great power of abstracting water.

Nitrogen pentoxide, or nitric anhydride, is an unstable body, and is not used in the arts, it is interesting as being the highest oxide of nitrogen, and the anhydride of nitric acid, uniting with great energy with water to form this acid.

OXYGEN.

Symbol—O. Atomic Weight—15.88. Density—15.9.

Oxygen is a colorless invisible gas, possessing neither taste nor smell. It exists in the free state in the atmosphere, of which it constitutes about one-fifth by bulk, whilst, in combination with the other elements, it forms nearly half the weight of the solid earth, and eight-ninths by weight of water. The birth of the modern science of chemistry may be dated from the

discovery of oxygen. Oxygen gas can be prepared from the air, but it is more easily obtained from many compounds which contain it in large quantities. Oxygen may be prepared by heating red mercury oxide. This substance is made up of 200 parts by weight of mercury, and sixteen parts of oxygen, when strongly heated, it is decomposed, and yields metallic mercury and oxygen gas. Oxygen can be more cheaply obtained by heating potassium chlorate, a white salt

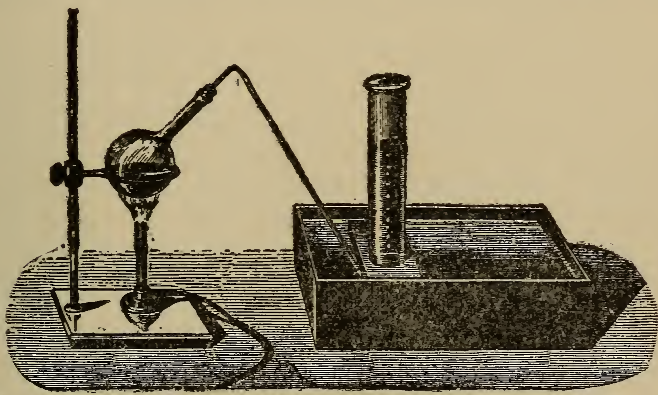

Fig. 34.

which yields on heating 39.2 per cent of its weight of this gas. In order to collect the oxygen thus given off, powdered potassium chlorate is placed in a small thin glass flask, furnished with a well-fitting cork, into which a bent tube is inserted. The lower end of the tube dips under the surface of water in a pneumatic trough, and the gas, on being evolved, bubbles out from the end of the tube, and is collected in jars or bottles filled with water, and placed with their mouths downwards in the trough. Fig. 34 shows the arrangement of the apparatus needed for the preparation of

oxygen gas. If a small quantity of manganese dioxide (black oxide of manganese) be mixed with the potassium chlorate, the oxygen is given off from the chlorate at a much lower temperature, and the evolution of the gas is facilitated, but the manganese dioxide undergoes no change whatever.

All the elements, with the single exception of fluorine, combine with oxygen to form oxides. In this act of combination, which is termed oxidation, heat is always, and light is frequently, given off. When bodies unite with oxygen, evolving light and heat, they are said to burn, or undergo combustion. All bodies which burn in the air burn with increased brilliancy in oxygen gas, and many substances, such as iron, which do not burn in the air, may be made to do so in oxygen. A redhot chip of wood, or a taper with glowing wick, is suddenly rekindled and bursts into flame when plunged into a jar of this gas. Sulphur, which in the air burns with a pale lambent flame, emits in oxygen a bright violet light, and a small piece of phosphorus, when inflamed and placed in oxygen, burns with a dazzling light. If the jars in which these experiments have been performed be afterwards examined, it is found that the substances produced by combustion in oxygen possess acid characters, they have the power of turning red certain vegetable blue coloring matters, such as litmus. A bundle of fine iron wire can be easily burnt in oxygen by tipping the end with burning sulphur, and then plunging the iron thus tipped into a jar of the gas, the oxide of iron, formed by the combustion, drops down in the molten state.

Many other substances may be employed for the preparation of oxygen, if large quantities of the gas

are needed, manganese dioxide may be heated to redness in an iron bottle, 100 parts by weight of the oxide yield 12.3 by weight of oxygen. Another interesting decomposition by which oxygen is set free is that effected by sunlight upon the carbonic acid gas contained in the air, this is accomplished by means of the green coloring matter of plants. Sunlight has the power, in presence of this green coloring matter, of decomposing carbonic acid, the carbon is taken up by the plant for its growth, while the oxygen is set free, and is afterwards used by animals for the support of the process of respiration. In the act of inspiration, animals breathe in the oxygen of the air, whilst in that of expiration, they breathe out carbonic acid gas. Hence oxygen is necessary to animal life, this gas was formerly termed vital air. The chemical change which oxygen effects upon the body of the animal is in fact identical with that which goes on when a piece of charcoal burns in the air or oxygen, this may be rendered evident by a simple experiment. If some clear lime-water be poured into a bottle of oxygen in which charcoal has been burnt, the lime-water will become milky, owing to the formation of a compound of lime and carbonic acid (called chalk), this acid being produced by the combustion, if the air contained in the lungs be next blown through a piece of glass tubing into some more clear lime-water, a turbidity (from the formation of chalk) will at once occur, proving that carbonic acid gas is given off from the lungs. This carbonic acid arises from the oxidation of the constituents of the body, and by this oxidation the heat of the body, which is greater than that of surrounding inanimate objects, is sustained. When this chemical process stops

the animal dies, and the temperature of the body sinks to that of the neighboring objects. Carbonic acid, nitrogen, and some other gases cause death when inhaled, because they do not contain free oxygen, and hence the process of oxidation in the body ceases. This cause of death is independent of any poisonous action of the gases.

OZONE.

Symbol—O_3. Molecular Weight—47.64. Density—23.8.

If a series of electric discharges be sent through a tube containing pure and dry oxygen, only a small portion of the gas is converted into ozone. If the ozone, however, is removed as soon as formed, by a solution of iodide of potassium, for example, the whole of the oxygen can be gradually converted into ozone. In order to obtain the maximum production of ozone, pure oxygen gas is allowed to pass through an apparatus, Fig. 35, which consists essentially of an iron tube **BB** turned very truly on the outside, through which a current of cold water can be passed by means of the tubes **CC**. Outside this metal cylinder is one of glass **AA** very slightly larger than the iron one. By means of tubes **DD** air or oxygen can be passed through the annular space between the two cylinders. Part of the outer cylinder at **G** is covered with tinfoil. The outer tinfoil coating and the inner metal cylinder are connected with the poles of an induction coil at **E** and **F**. By this means the oxygen is subjected to a series of silent discharges, by which it is converted partially into ozone. Oxygen gas through which an electric spark has been passed possesses a peculiar smell, and at once tarnishes a bright surface of mercury. This

peculiar strongly-smelling substance, which is called ozone, is capable of liberating iodine from potassium iodide, and of effecting many other oxidizing actions. Ozone is also produced in many other ways.

It is evolved at the positive pole in the electrolysis of acidulated water.

It is formed by the discharge from an electrical machine through air or through oxygen gas.

When fluorine is passed into water at 0°, the oxygen liberated contains 10 to 14 per cent of ozone.

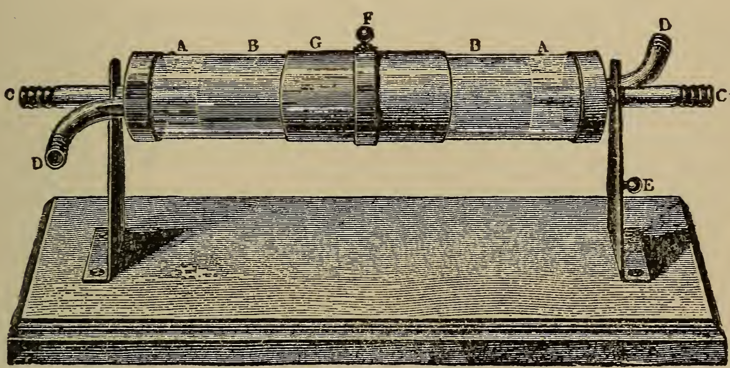

Fig. 35.

By acting with strong sulphuric acid upon dry barium dioxide, oxygen is given off which contains a considerable quantity of ozone.

When oxygen is passed over heated manganese dioxide, cobalt oxide or certain other metallic oxides, ozone is produced.

Ozone is obtained when potassium permanganate and sulphuric acid are distilled together in a vacuum.

It is stated that ozone is formed during combustion,

and can be recognized by its smell when a current of air is blown through the upper portion of a flame.

Powerful oxidizing agents are produced on the oxidation of phosphorus, turpentine and other oils by atmospheric oxygen, and these were formerly thought to be ozone.

PLATINUM.

Symbol—Pt. **Atomic Weight—193.4.**

Platinum is a comparatively rare metal, which always occurs in the native state, and generally alloyed with other metals. This alloy occurs in small grains in detritus and gravel in Siberia and Brazil, it has not been found in situ in the original rock, which probably belongs to the old plutonic series.

The original mode of obtaining the metal was to dissolve the ore in aqua regia, and precipitate the platinum with sal-ammoniac, as the insoluble double chloride of ammonium and platinum. This precipitate, on heating, yields metallic platinum in a finely divided or spongy state, and this sponge, if forcibly pressed and hammered when hot, gradually assumes a coherent metallic mass, the particles of platinum welding together, when hot, like iron. A new mode of preparing the metal has recently been employed, the ore being melted in a very powerful furnace. In this way a pure alloy of platinum, iridium, and rhodium is formed, the other constituents and impurities of the ore either being volatilized by the intense heat, or absorbed by the lime of which the crucible is composed. This alloy is in many respects more useful than pure platinum, being harder and less easily attacked by acids than the pure metal.

Platinum possesses a bright white color, and does not tarnish under any circumstances in the air, it is extremely infusible, and can only be melted by the heat of the oxy-hydrogen blowpipe. It is unacted upon by the ordinary acids, but dissolves in aqua regia, and hence platinum vessels are much used in the laboratory. Caustic alkalies, however, act upon the metal at high temperatures. When finely divided, metallic platinum has the power of condensing gases on to its surface in a remarkable degree.

Platinum and Oxygen Compounds.

Platinum and oxygen unite in two proportions to form Platinum monoxide, PtO, and Platinum dioxide, PtO_2. The first of these oxides is a black powder, easily decomposed on heating, and yielding a series of unstable salts, the second is obtained as a brown hydrate, by adding to a solution of platinic nitrate half its equivalent of caustic potash, the hydrate, when heated, first loses its water, forming the anhydrous oxide, and then parts with its oxygen, leaving the metal platinum.

Platinum and Chlorine Compounds.

Platinum dichloride: Formula—$PtCl_2$.

Platinum tetrachloride: Formula—$PtCl_4$.

Platinum dichloride is a green insoluble powder, obtained by heating the higher chloride to 200°. Platinum tetrachloride is the most important platinum compound. It is obtained as a yellowish-red solution by dissolving the metal in aqua regia, on evaporation, crystals of a compound of platinum tetrachloride with

hydrochloric acid separate out. Platinum tetrachloride combines with many alkaline chlorides to form double salts. These compounds with potassium and ammonium are insoluble in water, and are isomorphous, crystallizing in cubes, while the sodium salt is soluble, and crystallizes in large prisms.

POTASSIUM.

Symbol—K. **Atomic Weight—38.85.**

The metal potassium was discovered by the decomposition of the alkali potash into potassium, hydrogen, and oxygen, by means of a powerful galvanic current. Before this time the alkalies and alkaline earths were supposed to be elementary bodies. The metal is now prepared by heating together potash and carbon to a high temperature in an iron retort. The carbon, at the high temperature, is able to take the oxygen from the potash, forming carbon monoxide, which escapes as a gas, while the potassium, which is volatile at a red heat, distils over. The preparation of this metal is attended with many difficulties, and requires special precautions, as the vapor of potassium not only takes fire when brought in contact with the air, but decomposes water, combining with the oxygen and liberating hydrogen, hence the metallic vapor must be cooled by rock oil or naphtha, which contains no oxygen. The metal thus prepared must be distilled a second time, in order to purify it and free it from a black, explosive compound, which invariably forms in the original preparation.

Potassium, thus prepared, is a bright, silver-white metal, which can be easily cut with a knife at the or-

dinary atmosphere temperature, it is brittle at 0°, and melts at 62°.5, and does not become pasty before melting. When heated to a temperature below red heat, potassium sublimes, yielding a fine, green-colored vapor. This metal rapidly absorbs oxygen when exposed to the air, and gradually becomes converted into a white oxide. Thrown into water, one atom of potassium displaces one of hydrogen from the water, forming potassium hydroxide, or potash. This takes place with such force that the heat developed is sufficient to ignite the hydrogen thus set free, and the flame becomes tinged with the peculiar purple tint characteristic of the potassium compounds, while the water attains an alkaline reaction from the potash which is formed. Potassium also combines directly with chlorine and sulphur, and many other non-metals, evolving heat and light.

The original source of potassium compounds is the felspar of the granitic rocks of which the earth is composed, as these contain from two to three per cent of this metal. Up to the present time, this source has not been used for the manufacture of the potassium salts, as no cheap and easy mode has yet been available for separating the potash from the silicic acid, with which it is combined in felspar. Plants, however, are able slowly to separate out and assimilate the potash from these rocks and soils, so that, by burning the plant and extracting the ashes with water, soluble potassium salt is obtained. This is the crude potassium carbonate, called, when purified by re-crystallization, pearl-ash, and it is from this substance that a large number of the potassium compounds are obtained. Some of the other potassium salts, such as the nitrate and chloride, are

found in large quantities in various localities as deposits on the surface, or in the interior, of the earth. Potassium chloride occurs in beds, together with rock salt. Another inexhaustible source of potassium compounds is sea-water.

Potassium and Chlorine Compounds.

Potassium Chloride: Formula—KCl.

This salt occurs in certain saline deposits, and also exists in large quantities in sea-water. It crystallizes in cubes like sodium chloride, and is now much employed for the preparation of other potassium salts.

Potassium and Iodine Compounds.

Potassium Iodide: Formula—KI.

A very soluble salt, crystallizing in cubes, obtained by dissolving iodine in solution of caustic potash, and evaporating and igniting the solid mass to redness, this salt is much used as a medicine.

Potassium and Oxygen Compounds.

Potassium combines with oxygen in three proportions, forming three well-defined oxides of the formula, potassium monoxide, K_2O, potassium dioxide, K_2O_2, and potassium tetroxide, K_2O_4.

Potassium monoxide is obtained by allowing thin pieces of the metal to oxidize in dry air. It is a greyish-white, brittle substance, which melts a little above red heat, and volatilizes only at a very high temperature. This oxide combines with water with evolution of great heat, producing potassium hydroxide, or pot-

ash, from which water cannot again be separated by heat.

The dioxide and tetroxide are produced when potassium is oxidized at high temperatures.

Potassium, Oxygen and Hydrogen Compounds.

Potassium Hydroxide or Caustic Potash: Formula—KOH.

Caustic Potash is prepared by boiling one part of potassium carbonate with twelve parts of water, and adding slaked lime prepared from two-thirds part of quicklime. In this reaction calcium carbonate (chalk) is formed, which falls to the bottom as a heavy powder, caustic potash remaining in solution. The clear liquid, which should not effervesce on addition of an acid, is evaporated in a silver basin to dryness, fused by exposure to a stronger heat and cast into sticks in a metallic mould. Thus prepared caustic potash is a white substance, soluble in half its weight of water, and acts as a powerful cautery, destroying the skin. It is largely used in the arts and manufactures for soap-making, and is employed in the laboratory for various purposes.

Potassium, Carbon and Oxygen Compounds.

Potassium Carbonate: Formula—K_2CO_3.

This salt receives the commercial name of potash, or pearl-ash, and is exported in large quantities from Russia and America. The crude substance is prepared by boiling out the ashes of plants with water, and evaporating the solution to dryness, a pure salt may be afterwards obtained by separating the impurities by

crystallization. The leaves and small twigs of plants contain more potash than the stems and large branches. Potassium carbonate can be obtained perfectly pure by heating pure potassium tartrate to redness, and separating the carbonate formed by dissolving in water. This salt absorbs water from the air, or is deliquescent, and is, therefore, very soluble in water, it also turns red litmus blue, or possesses a strongly alkaline reaction.

Potassium, Chlorine and Oxygen Compounds.
Chlorate of Potash: Formula—$KClO_3$.

The action of chlorine on potash and the production of this salt is done on a large scale by decomposing calcium chlorate, made by saturating hot milk of lime with excess of chlorine, by means of potassium chloride. Potassium chlorate being but slightly soluble in cold water separates out in large tubular crystals, while the soluble calcium chloride remains dissolved. This salt is largely used in medical practice.

Potassium, Hydrogen, Carbon and Oxygen Compounds.
Bicarbonate of Potash: Formula—$KHCO_3$.

This substance is formed when a current of carbonic acid gas is passed through a strong solution of potassium carbonate. It may be considered as dibasic carbonic acid, in which one atom of hydrogen is replaced by one of potassium. It is a white salt, not so soluble as potassium carbonate, the solution is nearly neutral to test paper.

Potassium, Nitrogen and Oxygen Compounds.

Potassium Nitrate or Saltpeter: Formula—KNO_3.

This important salt occurs as an efflorescence on the soil of several dry tropical countries. It may be artificially prepared by the process of nitrification, in which animal matter (containing nitrogen) is exposed in heaps, mixed together with wood-ashes and lime to the action of the air, the organic matter gradually undergoes oxidation, nitric acid being formed, and this unites with the lime and the potash to form nitrates. The salt is obtained from both of these sources by boiling out the soil or deposit with water, adding potassium carbonate to decompose the nitrate of calcium, and allowing the nitre to crystallize out. It dissolves in seven parts of water at 15°, and in its own weight of hot water. It contains nearly half its weight of oxygen, with which it parts on heating with carbon or other combustible matter. For this reason, nitre is largely used in the manufacture of gunpowder and fireworks.

Potassium, Sulphur and Oxygen Compounds.

Sulphate of Potash: Formula—K_2SO_4.

Sulphate of Potash is contained in the ashes of both sea and land plants, and is only slightly soluble in water. A second sulphate termed hydrogen potassium sulphate, $HKSO_4$, or bisulphate of potash, is a soluble salt obtained in the process of the manufacture of nitric acid.

PHOSPHORUS.

Symbol—P. **Atomic Weight—30.77.**

Phosphorus does not occur free in nature, but is found in combination with oxygen and calcium in large quantities in the bodies, and especially the bones, of animals, in the seeds of plants, and also as the minerals phosphorite and apatite. When bones are burnt, a white solid mass is left behind, this is called calcium phosphate (phosphate of lime). Animals obtain the phosphate necessary for the formation of their tissues, from plants. Plants, again, draw their supply from the soil, whilst soils derive their phosphates from small quantities existing in the oldest granite rocks, by the disintegration of which the fertile soils have been produced. Phosphorus appears also to be a very necessary ingredient in the brain and other centers of the nervous action.

Phosphorus is prepared from powdered bone-ash, by mixing it with two-thirds of its weight of sulphuric acid and 15 to 20 parts of water. The sulphuric acid decomposes the bone-ash, forming calcium sulphate, or gypsum, which separates out as a white insoluble powder, while the greater part of the phosphorus in the bones comes into solution in combination with calcium, oxygen, and hydrogen, forming calcium hydrogen phosphate, a salt commonly known as superphosphate of lime. The liquid is drawn off clear, evaporated down to a syrup, and then mixed with powdered charcoal, dried, and heated to redness in an earthenware retort, the neck of which dips under water. Half the phosphorus is liberated together with carbon monoxide, and

distils over, collecting under the water in yellow drops, while the other half remains behind in the retort as calcium pyrophosphate.

In order to purify the phosphorus thus prepared, it may again be distilled, or pressed when melted under hot water through leather, it is then cast into sticks and kept under cold water. Phosphorus is an exceedingly inflammable and oxidizable substance, and requires great care in its preparation. It is manufactured on a very large scale for making the composition for the tips of matches. Phosphorus is a slightly yellow semi-transparent solid, resembling white wax both in appearance and consistency, but at low temperature it becomes brittle. It melts at 44°, forming a transparent liquid. It boils at 290°, giving rise to a colorless gas. In the air it gives off white fumes, emitting a pale phosphorescent light in the dark, whence its name, it is then undergoing a slow combustion, the white fumes consisting of phosphorus trioxide. At a temperature very little above its fusing point, phosphorus takes fire in the air, entering into active combustion, and forming phosphorus pentoxide, or phosphoric anhydride. The ignition of phosphorus takes place by slight friction, or by a blow, and even the heat of the hand may cause this substance to ignite, hence great care must be taken in handling phosphorus, and it should always be cut under water. Phosphorus does not dissolve in water, alcohol, or ether, but it is slightly soluble in oils, and very readily soluble in carbon disulphide.

If yellow phosphorus be exposed to a temperature of about 240° for some hours in an atmosphere incapable of acting chemically on it, such as hydrogen

or carbon dioxide, it is found to have undergone a very remarkable change, being wholly converted into a dark red opaque substance, altogether insoluble in carbon disulphide. The weight of red substance produced is exactly equal to that of yellow phosphorus used. This is called red, or amorphous phosphorus, and differs much in its properties from the yellow modification, especially in its inflammability, as it does not take fire in the air until heated to above 260°, when it becomes reconverted into the ordinary form, and burns with the formation of phosphorus pentoxide. The sudden conversion of yellow into red phosphorus can be shown by heating a small piece of ordinary phosphorus in a dry tube with a mere trace of iodine, combination at once occurs, a small trace of volatile phosphorus iodide is formed, and the remainder of the phosphorus is converted into the red modification. The red or amorphous modification of phosphorus can also be obtained in a crystallized form by heating red phosphorus in a tube with metallic lead. The phosphorus dissolves in the melted lead, and on cooling separates out in crystals, which possess a bright black metallic lustre.

Phosphorus appears luminous in the dark, when in contact with moist air, and it evolves fumes possessing a strong garlic-like smell. These fumes are poisonous, producing phosphorus-necrosis, a disease in which the bones of the jaw are destroyed, and one by which scrofulous subjects are the most easily affected. The luminosity of phosphorus in the air depends upon its slow oxidation, with formation of phosphorous acid. In this act of combination so much heat is evolved, that if a large piece of phosphorus be allowed to lie exposed to the air it at last melts and then takes fire. The lumi-

nosity and oxidation of phosphorus are best seen by pouring a few drops of the solution of this body in carbon bisulphide on to a piece of filter paper and allowing the solution to evaporate. In the dark the paper soon begins to exhibit a bright phosphorescence, and after a short time the phosphorus takes fire and burns. It was formerly believed that phosphorus becomes luminous in gases upon which it can exert no chemical action, such as hydrogen or nitrogen. This is not so, the luminosity which has been observed in these cases being due to the presence of traces of oxygen. From these facts it would naturally be inferred that phosphorus must be more luminous in pure oxygen than in air, this is not the case. At temperatures below 20° phosphorus is not luminous in pure moist oxygen, indeed it may be preserved for many weeks in this gas without undergoing the slightest oxidation. If, however, the gas be diluted by admixture with another indifferent gas, or if it be rarified, the phosphorescence is at once observed. The phenomenon can be very beautifully shown by placing a stick of phosphorus in a long tube **a**, Fig. 36, closed at one end and open at the other, and partly filled with mercury, into which some pure oxygen is brought. The open end of the tube is connected by a rubber tube with the vessel **b** containing mercury, so that, by raising or lowering the vessel, the pressure on the gas can be regulated. If the pressure be so arranged that it does not amount to more than one-fifth of an atmosphere, the phosphorus will be seen to be brightly luminous in the dark. If the pressure be then gradually increased, the light will become less and less distinct, until, when the level of the mercury is the same in both vessels, the luminosity has entirely

ceased. The phosphorescence can, however, at once be brought back again by lessening the pressure.

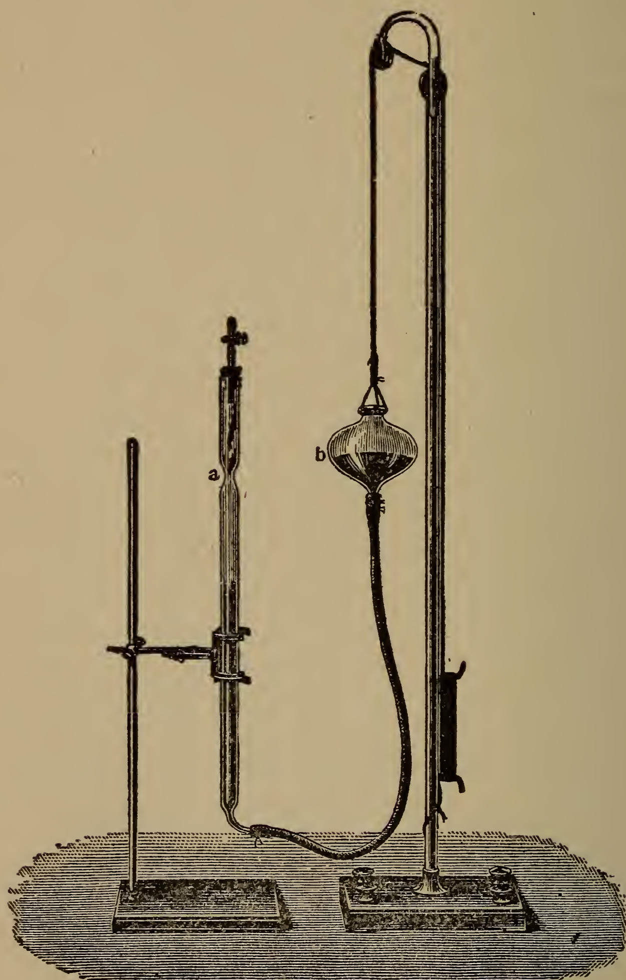

Fig. 36.

SILICON.

Symbol—Si. **Atomic Weight—28.2.**

Silicon, next to oxygen, is the most abundant element known. It does not occur, however, in the free state, but always combined with oxygen to form silicon dioxide, or silica. Silicon dioxide exists nearly pure in quartz or rock crystal, in flint, sand, and in a variety of minerals. Silicon also occurs combined with metals and oxygen, forming metallic silicates, and of these the greater part of almost all known rocks, especially the primary rocks, is composed.

In order to obtain silicon in the free state, a compound of this substance with fluorine and potassium, potassium silico-fluoride, is heated with metallic potassium, a violent reaction occurs, and when the contents of the tube in which the decomposition was effected are put into water, silicon is left undissolved in the form of a brown amorphous powder. Silicon can be obtained in three different modifications: Amorphous, graphitoidal, and crystalline. The graphite form of silicon is prepared by heating the brown amorphous powder to a high temperature, when the mass contracts, and becomes much more dense. Crystalline silicon is best obtained by fusing the mixture which gives brown silicon with zinc. On cooling the mass, crystals of silicon are found to be deposited on the zinc, which latter can easily be removed by solution in an acid. Silicon thus obtained is hard enough to scratch glass, it may be fused at a temperature between the melting point of cast-iron and steel.

Pure silicon may best be prepared by heating a mix-

ture of powdered quartz and magnesium, in the calculated quantities with a quarter of its weight of calcined magnesia in a fire-clay crucible, first at 300° to 400° in order thoroughly to dry the material, and then for a few minutes at a red heat, when a vigorous reaction takes place.

After the action has ceased, the cooled mass is treated with hydrochloric acid to dissolve out the magnesia, then warmed several times alternately with hydrofluoric and sulphuric acids to remove unchanged silica, finally washed with water and dried by heating in a stream of dry hydrogen. By employing precipitated silica and pure magnesium a very pure form of silicon may be obtained. An impure form, which may be advantageously used for the preparation of many of its derivatives, can be rapidly obtained by heating 40 grams of dry powdered white sand with 10 grams of magnesium powder in a wide test-tube. The reaction is a tolerably vigorous one, and if carried out with precipitated silica is accompanied by a brilliant flash of light.

Pure silicon prepared as above is an amorphous maroon-colored powder. When heated in air it oxidizes superficially, and it takes fire in oxygen at about 400°, burning brilliantly to silica, whilst it decomposes steam slowly at a dull red heat forming silica and hydrogen.

Amorphous silicon dissolves in many molten metals, uniting with the metal in some cases, as with iron, nickel, magnesium, and copper, to form metallic silicides, whilst in other cases, as, for example, with aluminium, zinc, and silver, no combination takes place, but on cooling the mass, silicon separates out in the crystalline form.

A different variety of amorphous silicon, which possesses remarkable reducing properties, is obtained when sparks are passed through liquid silico-ethane, the silicon, prepared in this way, reduces neutral solutions of potassium permanganate in the cold, and copper sulphate, gold and mercuric chlorides on boiling.

Silicon may also be obtained crystalline either in six-sided plates or long needles, and these forms are sometimes distinguished as graphitoidal and adamantine silicon. The crystalline modification of silicon is prepared by heating metallic aluminium to redness with three times its weight of potassium silico-fluoride in a fire-clay crucible, for about half-an-hour, when the following reaction takes place:

$$3K_2SiF_6+4Al=6KF+4AlF_3+3Si.$$

The temperature is then raised to about 1,000° and the silicon dissolves in the excess of aluminium, and on cooling crystallizes out in thin plates, which are separated from the aluminium and potassium fluorides by washing with hydrochloric and hydrofluoric acids.

Silicon and Carbon Compounds.

Silicon Carbide: Formula—SiC.

Silicon carbide is formed when a mixture of coke, sand and salt is fused in an electric furnace employing carbon terminals, and may be obtained pure by fusing silicon with the requisite amount of carbon in the electric furnace. It is also formed by the action of carbon on iron or calcium silicide, by fusing calcium carbide with silica in the electric furnace, and by the union of the vapors of carbon and silicon in the electric furnace, needles of the carbide being deposited. The

crystals are colorless or sapphire blue and are usually found as six-sided plates, but when formed by the combination of the two vapors they are prismatic. Silicon carbide is not oxidized by oxygen at 1,000° and is not attacked by sulphur vapor, by fused potassium nitrate, or by any acid. It is completely decomposed by chlorine at 1,200°, and on fusion with lead chromate it is gradually oxidized, while fused caustic potash slowly converts it into potassium carbonate and silicate.

The crude material, obtained by the method first described, is known as **carborundum** and is used as a cutting and polishing agent. It is manufactured in large quantities at Niagara Falls.

Silicon and Oxygen Compounds.

Silicon Dioxide or Silica: Formula—SiO_2. Molecular Weight—59.96.

Silicon forms only this oxide, which is an extremely important constituent of our planet. It is found not only in the mineral but also in the vegetable and animal kingdoms, existing in large quantity in the glassy straw of the cereals and of bamboos, and in the feathers of certain birds, which have been found to contain as much as 40 per cent of this substance, while silicic acid has been found in all forms of connective tissue. Vast deposits of pure silica in a very fine state of division occur in various parts of Germany, especially in Hanover and near Berlin. Large quantities of this substance are now used for a variety of purposes, especially for the preparation of dynamite, for filtering, and as a non-conducting medium for packing steam-pipes. In the mineral kingdom it is found in three distinct forms.

Silica melts in the oxy-hydrogen flame to a colorless glass which may be drawn out in threads. It slowly volatilizes when maintained at a high temperature in a furnace for some time, and can readily be made to boil when heated in an electric furnace.

When an alkaline solution of silica is heated in a sealed tube the glass is attacked and an acid silicate is formed from which silica separates out on cooling. If the temperature at which the deposition occurs be above 180°, the silica separates out as quartz, if below this point, it crystallizes out as tridymite, whilst at the ordinary temperature of the air it separates in the form of a hydrated amorphous mass.

The various forms of silica are employed in a variety of technical processes, especially its application to the manufacture of glass and porcelain. The colored varieties of silica are also largely used as gems and for other ornamental purposes. It is also possible to color the natural agates artificially. Thus brown or yellow agates or chalcedonies when strongly heated are changed into ruby carnelians, the yellow oxide of iron being thereby changed into the red anhydrous oxide. Many agates and chalcedonies are permeable to liquids, and in this way they may be artificially colored. This fact was known to the ancients and made use of in darkening the color of agates.

Silicon, Oxygen and Hydrogen Compounds.

Silicic Acid: Formula—$Si(OH)_4$.

If a solution of an alkali silicate, termed soluble glass, be acidified with hydrochloric acid, a portion of the silicic acid separates out as a gelatinous mass, whilst another portion remains in solution. If, on the

other hand, the solution is sufficiently dilute, no precipitate will occur, all the silicic acid remaining dissolved. This liquid contains silicic acid, hydrochloric acid, and common salt in solution. In order to separate the last two compounds from the first, the liquid is brought into a flat drum, the bottom of which consists of parchment paper, and this dialyser containing the liquid is allowed to swim on the surface of a large volume of water. The sodium chloride and the excess of hydrochloric acid pass through the membrane, whilst a clear aqueous solution of silicic acid remains behind.

This mode of separation was termed **dialysis** by its discoverer. Substances which crystallize, hence termed **crystalloids**, have as a rule the power of passing in solution through a membrane such as parchment or an animal membrane, whilst substances, such as gum and glue, which form jellies and are termed **colloids**, are unable to pass through such a diaphragm.

A fluid mixture of a colloid and a liquid is termed a **sol**, a watery mixture being called a **hydrosol**, whilst a firm mixture of a colloid and a liquid is named a "gel," thus silicic acid jelly is a **hydrogel** of silicic acid

These so-called colloidal solutions, which are thought by many to be suspensions and not true solutions, have the power of polarizing transmitted light.

By dialysis an aqueous solution of pure silicic acid may be obtained which contains 5 per cent of silica, and this may be concentrated by boiling it in a flask until it reaches a strength of 14 per cent. When heated in an open vessel, such as an evaporating basin, it is apt to gelatinize round the edge, after which the whole solidifies. The solution of silicic acid thus prepared has a feebly acid reaction and is colorless, limpid, and

tasteless. On standing for a few days the solution gelatinizes to a transparent jelly. This coagulation is retarded by the presence of a few drops of hydrochloric acid, or of caustic alkali, but it is brought about by even the smallest traces of an alkali carbonate.

If the clear solution be allowed to evaporate in a vacuum at 15° a transparent glass-like mass remains behind, which, when dried over sulphuric acid, possesses approximately the formula, $H_2SiO_3=SiO_2+H_2O$. This has been termed metasilicic acid, and an acid of the same composition has been obtained by dehydrating precipitated gelatinous silicic acid with 90 per cent alcohol.

SILVER.

Symbol—Ag. **Atomic Weight—107.13.**

Silver is found in the native state, as well as combined with sulphur, antimony, chlorine, and bromine. It is also contained in small quantities in galena, and it can be extracted with profit from the lead prepared from this ore, even when the lead contains only two or three ounces of silver to the ton. The method adopted for the extraction of the silver depends upon the fact that the whole of the silver can be concentrated into a small portion of lead, by crystallization, metallic lead free from silver separates out in crystals, and a rich alloy is left. When this reaches the concentration of 300 ounces of silver to the ton, the alloy undergoes the operation of cupellation, in which the mixture is melted in a furnace on a porous bed of bone-earth, and a blast of air blown over the surface. The lead oxidizes, and the oxide (litharge) fuses, and partly runs

away and partly sinks into the porous bed of the furnace, whilst the silver remains behind in the metallic state.

For the extraction of silver from the other ores, a process termed amalgamation is employed, in which mercury is used to dissolve the metallic silver. The argentiferous ores in which the silver occurs in combination with sulphur, are worked in a different manner. The ore is roasted in a furnace with common salt, by which means the silver sulphide is converted into chloride, the mixture is then placed in casks made to revolve, and scrap-iron and water are added. The iron reduces the silver to the metallic state, and when this is fully accomplished, metallic mercury is added. This forms a liquid amalgam with the silver, and gold, if any be present, and by distilling the mercury off, the silver is obtained in an impure state. Silver possesses a bright white color and a brilliant lustre, which it does not lose in pure air at any temperature, but when melted in the air it possesses the singular power of absorbing mechanically a large volume (twenty-two times its bulk) of oxygen, this gas it again gives out on solidifying, a phenomenon known as the spitting of silver.

Silver is probably the best conductor of heat and electricity known, and is extremely ductile, one gramme of metal can be drawn out into a wire of 2,600 metres in length. Sulphur combines at once with silver, forming a black sulphide, silver articles long exposed to the air tarnish from this cause. Silver is easily soluble in nitric acid, the nitrate being formed and nitric oxide gas being evolved.

Silver and Chlorine Compounds.

Silver Chloride: Formula—AgCl.

Of the insoluble salts, the Silver chloride is the most important. This salt occurs in nature, and is then known as horn silver, and is precipitated as a white curdy mass when solutions of a chloride and a silver salt are brought together. When exposed to sun or day-light, the white chloride becomes tinted a purple color, which increases in shade as the action of light continues. This coloration arises from a partial decomposition of the salt, a small quantity of argentous chloride and free hydrochloric acid being formed.

Silver and Oxygen Compounds.

Silver suboxide, Ag_4O, Silver monoxide, Ag_2O and Silver dioxide, Ag_2O_2.

Silver suboxide: Formula—Ag_4O.

Silver forms three compounds with oxygen. The first of these is called Silver suboxide, it is a black powder which readily undergoes decomposition.

Silver monoxide: Formula—Ag_2O.

The second, termed Silver monoxide, is obtained in the form of a brown precipitate, when caustic potash is added to a solution of silver nitrate. From this oxide, which is decomposed into metal and oxygen on heating, the ordinary silver salts may be derived by dissolving it in acids.

Silver dioxide: Formula—Ag_2O_2.

The third oxide is called Silver dioxide, and is obtained as a black powder by the action of ozone on metallic silver.

Silver, Nitrogen and Oxygen Compounds.

Nitrate of Silver: Formula—$AgNO_3$.

Silver nitrate is the most important soluble salt of silver. It is obtained in the form of large transparent tubular crystals on evaporating a solution of silver in nitric acid, and is soluble in its own weight of cold and half its weight of hot water, and in four parts of alcohol. Silver nitrate fuses easily on heating, and when cast into sticks goes by the name of lunar caustic. This salt undergoes decomposition when exposed to the sunlight in contact with organic matter, and a black substance, consisting of the suboxide, is formed. It is employed in the manufacture of an indelible ink for marking linen and other fabrics, and for photographic purposes.

SODIUM.

Symbol—Na. **Atomic Weight—22.87.**

This metal was discovered by the decomposition of soda with the galvanic current. It can be procured more easily than potassium by reducing the carbonate in presence of carbon, and is now manufactured in large quantities for the preparation of other metals, especially magnesium and aluminium. The apparatus employed for the preparation of this metal is the same as that used for potassium. The metal distills over

when condensed, and drops into rock oil. Sodium is a silver-white metal, soft at ordinary temperatures, and melting at 95.6°, it volatilizes below a red heat, yielding a colorless vapor. When thrown upon water it floats, and rapidly decomposes the water with disengagement of hydrogen, soda being formed. If the water be hot or be thickened with starch, the globule of metal becomes so much heated as to enable the hydrogen to take fire. The compounds of sodium are very widely diffused, being contained in every particle of dust. They exist in enormous quantities in the primitive granitic rocks, but they are most readily obtained from sea-water, which contains nearly three per cent of sodium chloride (common or sea salt). Sodium carbonate was formerly obtained from the ashes of sea-plants or kelp, as potassium carbonate is still prepared from the ashes of land plants, but at present the sodium carbonate is altogether manufactured, on a large scale, from sea-salt.

Sodium, Carbon and Oxygen Compounds.

Sodium Carbonate or Soda-ash: Formula—Na_2CO_3.

This substance, known in commerce as soda-ash, is manufactured on an enormous scale, and used for glass-making, soap-making, bleaching, and various other purposes in the arts. Formerly it was prepared from barilla or the ashes of sea-plants, but now it is wholly obtained from sea-salt by a series of chemical decompositions and processes.

Sodium and Chlorine Compounds.

Chloride of Sodium or Common Salt. Formula—NaCl.

It is from this salt that almost all the other sodium compounds are prepared. Sodium chloride occurs in thick beds in various parts of the world. It is likewise prepared from sea-water by evaporation or by freezing, and from certain brine springs by evaporation. When slowly deposited sodium chloride crystallizes in regular cubes. It is soluble in about two and a half parts of water at 15°, and does not dissolve sensibly more in hot than in cold water.

Sodium and Oxygen Compounds.

Sodium Oxide, Na_2O and Sodium Dioxide, Na_2O_2.

Sodium Oxide: Formula—Na_2O.

Sodium Oxide is formed when sodium is oxidized in dry air or oxygen at a low temperature, a white powder being formed. This takes up moisture with great avidity, forming sodium hydroxide, or soda, from which water cannot again be separated by heat alone, but which can again be converted into the oxide by heating with sodium.

Sodium Dioxide: Formula—Na_2O_2.

Sodium Dioxide is a yellowish-white powder, which is formed when sodium is heated in oxygen to 200° C. It is soluble in water, but the solution readily decomposes, giving off oxygen and leaving sodium hydroxide.

Sodium, Hydrogen and Oxygen Compounds.

Sodium Hydroxide or Caustic Soda: Formula—NaHO.

Sodium Hydroxide is a white solid substance, fusible below a red heat, and less volatile than the corresponding potassium compound. It is very soluble in water, acts as a caustic, is powerfully alkaline and is largely used in soap-making. The manufacture of solid caustic soda is carried on on a large scale, by boiling lime and sodium carbonate together with water, and evaporating down the clear solution.

Sodium, Hydrogen, Carbon and Oxygen Compounds.

Sodium Bicarbonate: Formula—$NaHCO_3$.

Bicarbonate of Soda is obtained by exposing the crystallized carbonate in an atmosphere of carbonic acid gas. It is a white crystalline powder, which on heating is readily converted into sodium carbonate. The bicarbonate is chiefly used in medicine, and for the production of effervescing drinks.

If sodium carbonate be added to a solution of trihydrogen phosphate, effervescence will at once ensue from the liberation of carbonic acid, and if the carbonate be added until the solution ceases to redden litmus paper, a salt will be obtained on evaporation which crystallizes in large transparent prisms. This is common sodium phosphate, its composition is represented by the symbol Na_2HPO_4, with twelve molecules of water of crystallization. If caustic soda be added to a solution of this common phosphate, a salt termed the subphosphate, crystallizes out in small needles on evaporation, the composition of this salt is Na_3PO_4,

with twelve atoms of water of crystallization. And if phosphoric acid be added to a solution of common phosphate, the so-called sodium superphosphate is formed, NaH_2PO_4.

Sodium, Nitrogen and Oxygen Compounds.

Sodium Nitrate or Saltpeter: Formula—$NaNO_3$.

Sodium nitrate is found in large beds in Peru and Chili, and is termed Chili saltpeter. It is imported in large quantities and used as a manure, and also in the preparation of nitric acid, being cheaper than nitre. For this latter purpose a hot concentrated solution of this salt is mixed with a hot saturated solution of potassium chloride, on cooling, potassium nitrate separates out in crystals, and sodium chloride remains in solution.

Sodium, Sulphur and Oxygen Compounds.

Sulphate of Soda or Glauber's Salts: Formula— $Na_2SO_4+10H_2O$.

Sodium sulphate is known in commerce as Glauber's salts, and in the anhydrous state as salt-cake. It occurs in the water of many mineral springs, and is used in medicine, while as salt-cake it is employed in large quantities in glass manufacturing.

SULPHUR.

Symbol—S. **Atomic Weight—31.8.**

Sulphur occurs both free and combined in nature. It is found free in certain volcanic countries, especially in Sicily, Iceland and Mexico, and occurs crystallized in yellow transparent crystals. It exists in combina-

tion with many metals, forming compounds termed sulphides, which constitute the common ores from which the metals are usually obtained, thus lead sulphide, or Galena, zinc sulphide, or Blende and copper sulphide, are the substances from which those metals are generally procured. Sulphur also is found in nature combined with metals and oxygen, to form a class of salts called sulphates. Of these, calcium sulphate or gypsum, barium sulphate or heavy spar, sodium sulphate or Glauber's salt, occur in the largest quantity. Sulphur likewise occurs combined with hydrogen, as a gas called Sulphuretted Hydrogen, H_2S, in the waters of certain springs. In order to obtain pure sulphur, the mineral containing the crude substance mixed with earthy impurities is heated in earthenware pots, the sulphur distills over in the form of vapor, which is condensed in similar pots placed outside the furnace. When brought to this country, the sulphur thus obtained is refined or purified by subjecting it to a second distillation. If the vapor of sulphur is quickly cooled below its melting point, it solidifies in the form of a fine crystalline powder called Flowers of Sulphur, exactly as aqueous vapor, when cooled down below the freezing point of water, deposits as snow. When sulphur is gently heated, it melts, and may be cast into sticks, and is then known as brimstone or roll sulphur.

Sulphur exists in three modifications. The first is that in which sulphur crystallizes in nature, and the other two are obtained by melting sulphur. If melted sulphur be allowed to cool slowly, it crystallizes in long, transparent, needle-shaped, prismatic crystals, which are quite different in form from the natural

crystals of sulphur. These transparent crystals become opaque after exposure to the air for a few days, owing to each crystal splitting up into several crystals of the natural or permanent form. The third allotropic modification of sulphur is obtained by pouring melted sulphur heated to 230° into cold water. The sulphur thus forms a soft tenacious mass resembling rubber. This form of sulphur is, however, not permanent, in a few hours, at the temperature of the air, the mass assumes the ordinary brittle form of the element, while, if heated to 100°, it instantly changes to the brittle form, and thereby evolves so much heat as to raise its temperature up to 111°. These peculiar modifications become apparent when sulphur is heated. Sulphur melts, to begin with, at 115°, forms an amber-colored mobile liquid, as the temperature rises, the liquid becomes dark-colored, and attains the consistency of thick molasses, so that at about 230° it can scarcely be poured out of the vessel, heated above 250°, it again becomes fluid, and remains as a dark reddish-black colored thin liquid, until the temperature rises to 490°, when it begins to boil, and gives off a red-colored vapor.

Sulphur is an inflammable substance, and when heated in the air or in oxygen burns with a bluish flame, combining with the oxygen to form sulphur dioxide, often called sulphurous acid, which is given off as a gas, possessing the peculiar and well-known suffocating smell which is evolved when a common match is burnt. Sulphur combines directly with chlorine, carbon, and most other elements, while many metals burn in sulphur vapor as they do in oxygen gas, uniting to form sulphides. Sulphur is insoluble in water and most organic liquids, but both the natural variety and the

other crystalline or prismatic variety dissolve freely in carbon disulphide, while the tenacious form of sulphur is insoluble in this liquid. When deposited from solution in carbon disulphide, sulphur crystallizes in the ordinary natural form.

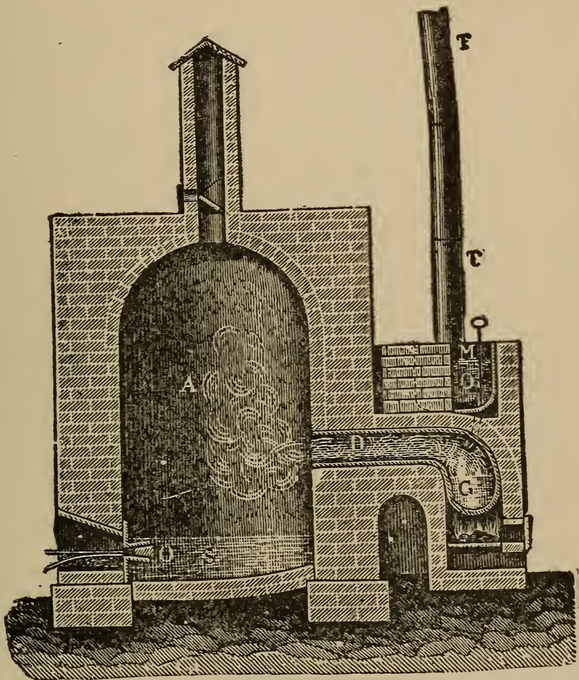

Fig. 37.

Commercial sulphur contains about 3 per cent of earthy impurities which can be removed by distillation, an arrangement for this purpose being shown in Fig. 37. The sulphur is melted in an iron pot **M** and runs from this by means of a tube into the iron retort **G** where it is heated to the boiling point, the vapor of

the sulphur then passes into the large chamber **A** which has a capacity of 200 cubic metres. In this chamber the sulphur is condensed, to begin with, in the form of a light yellow powder termed flowers of sulphur, just as aqueous vapor falls as snow when the temperature suddenly sinks below 0°. After a time the chamber becomes heated above the melting point of sulphur, and then it collects as a liquid which can be drawn off by means of the opening **O**. It is then cast in slightly conical wooden moulds, and is known as roll sulphur or brimstone. It is frequently also allowed to cool in the chamber and then obtained in large crystalline masses, known to the trade as block sulphur.

Another and much more important source from which sulphur is now obtained is the residue or waste in the soda manufacture, this consists of calcium sulphide mixed with chalk, lime, and alkali sulphides. The sulphur which this material contains was formerly altogether wasted, now it is economically regained in the alkali works on a very large scale. For this purpose the waste, which consists mainly of calcium sulphide, CaS, is treated in the presence of water with carbonic acid gas, calcium carbonate being thus produced whilst sulphuretted hydrogen is evolved.

The gas thus obtained is then mixed with a quantity of air sufficient to supply the amount of oxygen required and the mixture passed through a kiln, where it comes in contact with heated ferric oxide and the above reaction takes place, the sulphur formed being condensed in cooling chambers, and recovered in the pure state. A less pure sulphur is obtained in the same way from the sulphuretted hydrogen given off

during the distillation of the ammoniacal liquor from gas works.

The simplest mode of detecting sulphur in a compound is to mix the substance with pure sodium carbonate, and fuse it before the blowpipe on charcoal, or, to avoid the introduction of sulphur from the gas flame, to mix the substance with sodium carbonate and charcoal, and heat in a small closed crucible. Sodium sulphide is thus formed, and this may then be recognized by bringing the fused mass on to a silver coin and adding water. The smallest quantity of sulphur can thus be recognized by the formation of a brown stain of silver sulphide. Sulphur is almost always quantitatively determined as barium sulphate. If the body is a sulphide, as for instance, pyrites, it is finely powdered, and either fused with a mixture of sodium carbonate and nitre, the fused mass dissolved in water, and the filtrate, after acidifying by hydrochloric acid, precipitated with barium chloride, or the sulphide is oxidized with a mixture of nitric and hydrochloric acids or fuming nitric acid, the excess of acid removed by evaporation and barium chloride added, whereby insoluble barium sulphate is formed, and this, after washing and drying, is ignited and weighed. A further method is to fuse the mineral with 2 parts of caustic soda and 4 parts of sodium peroxide, the melt being then acidified and precipitated with barium chloride as above.

Sulphur and Oxygen Compounds.

Sulphur Dioxide, SO_2 and Sulphur Trioxide, SO_3.

Sulphur Dioxide: Formula—SO_2. Molecular Weight —63.59. Density—31.8.

Sulphur dioxide is formed not only by the combustion of sulphur, but also by the action of certain metals, such as copper, mercury, or silver, on concentrated sulphuric acid.

Sulphur dioxide is easily prepared for laboratory use. For this purpose a flask is half filled with copper turnings or fine copper foil, and so much strong sulphuric acid poured in that the copper is not quite covered. The mixture is next heated until the evolution of gas commences, the lamp must then be removed, as otherwise the reaction may easily become too violent, and the liquid froth over.

Pure sulphur dioxide is also produced when sulphur and sulphuric acid are heated together, thus

$$S+2H_2SO_4=3SO_2+2H_2O.$$

It is also formed by the decomposition of a sulphite, such as commercial sodium sulphite, which, when treated with warm dilute sulphuric acid, easily evolves the gas, thus

$$Na_2SO_3+H_2SO_4=Na_2SO_4+H_2O+SO_2.$$

A convenient method of employing this decomposition is to allow concentrated sulphuric acid to drop into a saturated solution of sodium bisulphite.

Sulphur dioxide is made on a large scale for the preparation of the sulphites, especially of sodium sulphite and calcium sulphite, which are obtained by passing the gas either into a solution of caustic soda or into milk of lime. For this purpose charcoal is heated together with sulphuric acid, when carbon di-

oxide is evolved, together with sulphur dioxide, but the presence of the former compound for the purpose above mentioned is not detrimental, thus

$$C+2H_2SO_4=2H_2O+CO_2+2SO_2.$$

The sulphur dioxide evolved in the roasting of certain metallic ores, which was formerly allowed to pass off into the atmosphere, is now frequently utilized for the preparation of the sulphites.

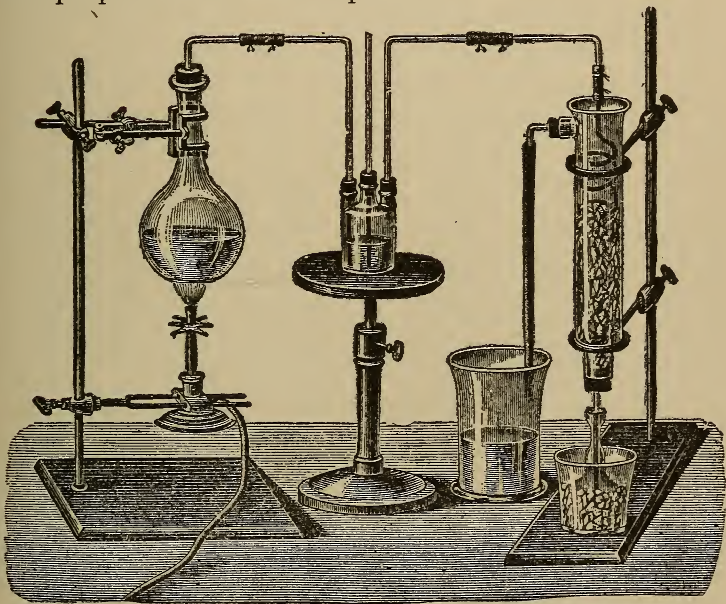

Fig. 38.

Sulphur dioxide is used in enormous quantities for the manufacture of sulphuric acid. For this purpose it is chiefly obtained by roasting pyrites. When, however, especially pure sulphuric acid is needed, the dioxide is prepared by burning pure sulphur.

Sulphur dioxide is a colorless gas, which occurs in

nature in certain volcanic emanations, as well as in solution in volcanic springs. It possesses the well-known suffocating smell of burning sulphur. It can be collected by downward displacement, like chlorine.

In order to prepare liquid sulphur dioxide in larger quantity, the apparatus Fig. 38 is used. The gas evolved by the action of sulphuric acid on copper is purified by passing through the wash-bottle, and afterwards passes through the spiral glass tube, surrounded by a freezing mixture of ice and salt. The liquid which condenses and falls into the flask placed beneath may be preserved by sealing the flask hermetically where the neck has been drawn out.

Sulphur Trioxide: Formula—SO_3. Molecular Weight—79.47. Density—39.73.

Sulphur trioxide is obtained by passing a perfectly dry mixture of sulphur dioxide and oxygen through a tube containing heated platinum sponge or platinized asbestos as in Fig. 39. The sulphur dioxide is evolved in the flask **a** and is mixed in the wash-bottle, which contains strong sulphuric acid, with the oxygen from a gas-holder coming in through the tube **b**. The mixture next passes through the cylinder **e** containing pumice-stone soaked in strong sulphuric acid in order to remove every trace of moisture, and then passes at **c** over the platinized asbestos. As long as this is not heated no change is observed, so soon, however, as it is gently ignited dense white fumes of the trioxide are formed which condense in a receiver **d**, cooled by a freezing mixture, in the form of long white needles. In order to obtain these crystals, every portion of the

apparatus must be absolutely dry, if even a trace of moisture be present the needles disappear at once, liquid sulphuric acid being formed. Instead of platinum, certain metallic oxides, such as copper oxide, ferric oxide, and chromic oxide, may be used.

A much more convenient method of preparing sulphur trioxide for experimental purposes is by the dis-

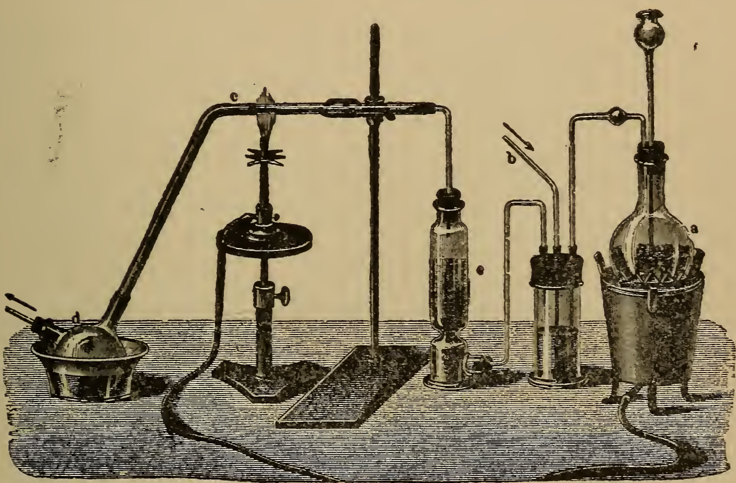

Fig. 39.

tillation of fuming sulphuric acid, which consists of a solution of the trioxide in sulphuric acid, and is manufactured in large quantity.

The trioxide absorbs moisture rapidly from the atmosphere, and evolves dense white fumes on exposure to the air, and when brought in contact with anhydrous baryta combines with it to form barium sulphate, $BaSO_4$, with such energy that the mass becomes incandescent.

Sulphur trioxide is now manufactured in very large quantities by the combination of sulphur dioxide and oxygen in presence of catalytic agents, such as platinum or ferric oxide.

TIN.

Symbol—Sn. **Atomic Weight—118.2.**

The ores of tin, although this metal has been known from very early times, occur in but few localities, and the metallic tin is not found in nature. The chief European sources of tin are the Cornish mines, where it is found as tin dioxide or tinstone. Tinstone is also met with in Malacca, and Borneo, and Mexico. In order to prepare the metal, the tinstone is crushed and washed, to remove mechanically the lighter portions of rock with which it is mixed, and the purified ore is then placed in a reverberatory furnace with anthracite or charcoal and a small quantity of lime. The oxide is thus reduced, and the liquid metal, together with the slag, consisting of silicate of lime, falls to the lower part of the furnace. The blocks of tin, still impure, are then refined by gradually melting out the pure tin, leaving an impure alloy behind. English tin generally contains traces of arsenic, copper, and other metals, that imported from Banca is nearly chemically pure.

Tin possesses a white color resembling that of silver, it is soft, malleable, and ductile, but possesses little tenacity, a wire two mms. in diameter breaking with a weight of sixteen kilos. When bent, pure tin emits a peculiar crackling sound. Tin melts at 235°, and is not sensibly volatile. Tin does not lose its lustre on exposure to the air, whether dry or moist, at ordinary

temperature, but if strongly heated it takes fire, and a white powder of stannic oxide is formed. Hydrochloric acid dissolves tin with the evolution of hydrogen and the formation of stannous chloride. Nitric acid also attacks the metal with great energy, nitrous fumes being given off and stannic oxide being left as a white powder.

Tin can easily be distinguished in solution by the formation of a splendid purple color called purple of cassius, formed when gold chloride is added to a dilute solution of stannous chloride. Tin is also easily reduced before the blowpipe in the form of white malleable beads. Tin is largely used in the arts for covering and thus protecting iron plates, or for tin-plating.

Tin and Chlorine Compounds.

Stannic Chloride: Formula—$SnCl_4$.

Stannic chloride is obtained by passing chlorine gas over metallic tin, it is a colorless liquid, boiling at 120° Centigrade, and having a vapor density of 9.2. It fumes strongly in the air, and forms a crystalline hydrate, when a small quantity of water is added, which easily dissolves in an excess. Stannic chloride is also used by dyers, and is prepared for this purpose by dissolving tin in cold nitro-hydrochloric acid.

Stannous Chloride: Formula—$SnCl_2$.

Stannous chloride is obtained by dissolving tin in hydrochloric acid, and separates out in needle-shaped crystals, when the solution is concentrated. Stannous chloride is termed **tin salts** in commerce, it is largely

manufactured for the calico-printer and dyer, who use it as a **mordant.**

Tin and Oxygen Compounds.

Tin Dioxide, SnO_2 and Tin Monoxide, SnO.

Stannic or Tin Dioxide: Formula—SnO_2.

Tin dioxide occurs native as tinstone, and it can be prepared as a hydrate in two conditions, possessing totally different properties. If tin be oxidized by nitric acid, hydrated stannic oxide is produced as a white powder insoluble in acids. On the other hand if to a solution of stannic chloride an alkali be added, a white precipitate is formed of hydrated stannic oxide, which is readily soluble in acids. Both of these varieties of hydrated stannic oxide form salts, the insoluble compound having been termed metastannic, and the soluble compound stannic acid. Sodium stannate is formed by boiling stannic oxide with soda and is largely used in calico-printing as a **mordant.**

Stannous or Tin Monoxide: Formula—SnO.

This is a black powder prepared by heating the stannous hydrate in an atmosphere of carbonic acid, it rapidly absorbs oxygen from the air, passing into stannic oxide. The hydrate falls as a white powder when a solution of a stannous salt is added to an alkaline carbonate.

Tin and Sulphur Compounds.

Stannous Sulphide: Formula—SnS.

Stannic Sulphide: Formula—SnS_2.

Of the sulphides of tin, stannous sulphide and stannic sulphide are the most important. The former is blackish-grey, and the latter a bright yellow crystalline powder known as mosaic gold is soluble in alkaline sulphides.

TUNGSTEN.

Symbol—W. **Atomic Weight—182.7.**

This metal occurs in tolerably large quantities combined with ferrous oxide in the mineral wolfram, and also with lime as scheelite. The metal has only been obtained as a greyish-black powder. Tungsten is employed in the arts, the addition of a small quantity imparts a great degree of hardness and other valuable qualities to steel.

Tungsten and Oxygen Compounds.

Two oxides of tungsten are known, Tungsten dioxide, WO_2, and Tungsten trioxide, WO_3. The former of these is obtained as a brown powder by heating the trioxide in an atmosphere of hydrogen, the latter, sometimes called tungstic acid, is obtained as an insoluble yellow powder by heating the native calcium tungstate with nitric acid. Tungsten trioxide forms a variety of somewhat complicated salts. The sodium compound is soluble, and has been used to add to the starch employed to stiffen light fabrics, the tungstate rendering the fabric uninflammable.

ZINC.

Symbol—Zn. **Atomic Weight—64.9.**

Zinc is an abundant and useful metal, closely resembling magnesium in its chemical characters, but it is much more easily extracted from its ores than this latter metal. The chief ores of zinc are the sulphide or blende, the carbonate or calamine, and the red oxide. In order to extract the metal, the powdered ore is roasted, or exposed to air at a high temperature, so as to convert the sulphide or carbonate into oxide, the roasted ore is then mixed with fine coal or charcoal and strongly heated in crucibles or retorts of peculiar shape, the zinc oxide is reduced by the carbon, carbon monoxide comes off, and the metallic zinc distills over, and is easily condensed.

Zinc is a bluish-white metal, exhibiting crystalline structure, it is brittle at the ordinary temperature, but when heated to about 130°, it may be rolled out or hammered with ease, whilst if more strongly heated to 200°, it is again brittle, and may be broken up in a mortar. Zinc melts at 423°, and at a bright red heat it begins to boil, and volatilizes, or if air be present it takes fire and burns with a luminous greenish flame, forming zinc oxide. Zinc is not acted upon by moist or dry air, and hence it is largely used in the form of sheets, and is employed as a protecting covering for iron, which when thus coated is said to be galvanized. Zinc easily dissolves in dilute acids with evolution of hydrogen, and it is thus used as the oxidizable portion of the galvanic battery. Brass is a useful alloy of one part of zinc and two of copper, German silver is an alloy of zinc, nickel, and copper.

Zinc and Oxygen Compounds.

Zinc Oxide: Formula—ZnO.

Zinc oxide is the only known compound of this metal with oxygen, and is obtained by burning the metal, or by precipitating a soluble zinc salt with an alkali, and heating the precipitate. Zinc oxide is an insoluble white amorphous powder, which when heated becomes yellow, but loses this color on cooling, it dissolves easily in acids, giving rise to the zinc salts.

The salts of zinc can be distinguished by the solubility of the oxide in excess of both potash and ammonia, by the white sulphide insoluble in acetic acid, and by the green color which a solution of cobalt chloride imparts to zinc salts when heated in the blowpipe.

Zinc, Carbon and Oxygen Compounds.

Carbonate of Zinc: Formula—$ZnCO_3$.

Zinc carbonate is an insoluble substance, occurring native as calamine, it cannot be prepared by precipitating a solution of zinc salt by an alkaline carbonate, as a quantity of oxide is precipitated along with the carbonate.

Chloride of Zinc: Formula—$ZnCl_2$.

Zinc chloride is a white soluble deliquescent substance, formed by burning zinc in chlorine, or by dissolving the metal in hydrochloric acid.

Sulphate of Zinc: Formula—$ZnSO_4$.

Zinc sulphate is a soluble salt, crystalizing in long prisms, and commonly called white vitriol. This salt is isomorphous with magnesium sulphate, and, like the latter salt, it forms a series of double salts with alkaline sulphates.

Zinc Sulphide: Formula—ZnS.

Zinc sulphide occurs as a crystalline mineral called blende, generally colored, from presence of iron and other impurities. It is obtained artificially as a white gelatinous precipitate, insoluble in acetic, but soluble in a mineral acid, formed when an alkaline sulphide is added to a soluble zinc salt.

DETERMINATION BY WEIGHT.

As it is the aim of the chemist to examine the properties of the elements and their compounds, and as the weight-determination of a substance is of the greatest importance, it becomes necessary for him to ascertain with great precision the proportion by weight in which these several elements combine, as well as that in which any one of them occurs in a given compound, and for this purpose the Balance is employed. By means of this instrument the weight of a given substance is compared with the unit of weight. It consists essentially of a light but rigid brass beam (Fig. 40), suspended on a fixed horizontal axis situated at its centre; and this beam is so hung as to assume a horizontal position when unloaded. At each end of the beam scale-pans are hung, one to receive the body to be weighed and the other for the weights. When each pan is equally weighted the beam must still retain its horizontal position or oscillate about this position, but when one pan is more heavily weighted than the other, the beam will incline on the side of the heavier pan. The balance is, therefore, a lever with equal arms, and it is evident that the weight of the substance relative to the unit weight employed is the sum of the weights necessary to bring the balance into equilibrium. The two important requisites in a balance are accuracy and sensibility, and these can only be gained by careful construction. It needs but little consideration to see that in a delicate balance the friction of the various parts must be re-

duced to a minimum. This is usually accomplished by suspending the beam by means of an agate knife-edge, working on an agate plane, whilst the pans are attached to each end of the beam by a somewhat similar arrangement. The position of the axis of suspension relatively to the centre of gravity of the beam is likewise a matter of consequence. If the axis of suspension

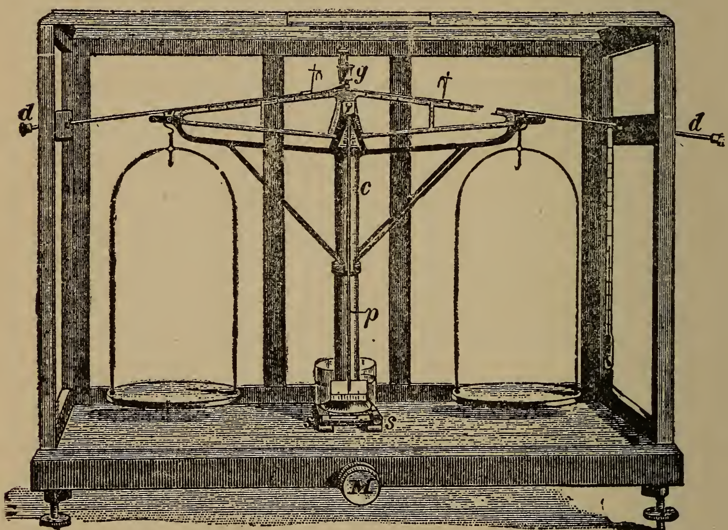

Fig. 40.

and the centre of gravity in a balance were coincident, the beam would remain stationary in all positions in which it might be placed. If the axis of suspension be placed below the centre of gravity the beam would be in a condition of unstable equilibrium. Hence the only case in which the balance can be used is that in which the point or axis of suspension is above the centre of gravity, for in this case alone will the beam

return to a horizontal position after making an oscillation, and in this case the balance may be considered as a pendulum, the whole weight of the beam and pans being regarded as concentrated at the centre of gravity. In order that the weight of the substance and the sum of the measuring weights in the scale-pan may be equal, it is evident that the axis of suspension must be exactly in the centre of the beam, or in other words, that the balance must have arms of equal length. It is also necessary that the balance should have great sensibility; that is, that it may be moved by the smallest possible weight; for this end it is likewise requisite that the vertical distance of the centre of gravity below the axis of suspension should be as small as possible. As the whole weight of the instrument may be regarded as concentrated at the center of gravity, it evidently requires a less force to act at the end of the beam to move the instrument when the distance of the centre of gravity from the point of suspension of the balance is small, than when that distance is greater, inasmuch as in the latter case the weight has to be lifted through a longer arc. The sensibilty of the balance is also increased, both by increasing the length of the beam and by diminishing the weight of the beam and of the load. When, however, the beam is made either too long or too light it cases to be rigid, and a serious source of error is introduced.

In all weighings with delicate balances it is necessary to have recourse to the method of weighing by vibration by which the excursions of the moving beam are accurately observed instead of its approach to the horizontal position.

A good chemical balance, such as is commonly used

for analytical work will indicate 0.0001 gram, when loaded with 50 to 100 grams in each pan. With a specially constructed balance by combining this method of vibrations with that of double weighing which consists in reversing the position of the loads in the two pans, it is possible with a load of 1 kilogram in each pan to ascertain definitely a difference of weight of 0.0001 gram or the 1/10,000,000 of the weight in either pan.

Comparison of the Metrical with the Common Measures.

MEASURES OF LENGTH.

	In English Inches.	In English Feet equals 12 Inches.	In English Yards equals 3 Feet.	In English Fathoms equals 6 Feet.	In English Miles equals 1,760 Yards.
Millimetre	0.03937	0.0032809	0.0010936	0.0005468	0.0000006
Centimetre	0.39371	0.0328090	0.0109363	0.0054682	0.0000062
Decimetre	3.93708	0.3280899	0.1093633	0.0546816	0.0000621
Metre	39.37079	3.2808992	1.0936331	0.5468165	0.0006214
Decametre	393.70790	32.8089920	10.9363310	5.4681655	0.0062138
Hectometre	3937.07900	328.0899200	109.3633100	54.6816550	0.0621382
Kilometre	39370.79000	3280.8992000	1093.6331000	546.8165500	0.6213824
Myriometre	393707.90000	32808.9920000	10936.3310000	5468.1655000	6.2138244

1 Inch equals 2.539954 Centimetres.
1 Foot equals 3.0479449 Decimetres.

1 Yard equals 0.91438348 Metre.
1 Mile equals 1.6093149 Kilometre.

MEASURES OF SURFACE

	In English Square Feet.	In English Square Yards equals 9 Square Feet.	In English Poles equals 272.25 Square Feet.	In English Roods equals 10,890 Square Feet.	In English Acres equals 43,560 Square Feet.
Centiare or sq. metre	10.7642993	1.1960333	0.0395383	0.000988457	0.0002471143
Are or 100 sq. metres	1076.4299342	119.6033260	3.9538290	0.098845724	0.0247114310
Hectare or 10,000 sq metres	107642.9934183	11960.3326020	395.3828959	9.884572398	2.4711430996

1 Square Inch equals 6.4513669 Square Centimetres.
1 Square Foot equals 9.2899683 Square Decimetres.

1 Square Yard equals 0.83609715 Square Metre or Centaire.
1 Acre equals 0.404671021 Hectare.

Comparison of the Metrical with the Common Measures.—Cont.

MEASURES OF CAPACITY.

	In Cubic Inches.	In Cubic Feet equals 1,728 Cubic Inches.	In Pints equals 34.65923 Cubic Inches.	In Gallons equals 8 Pints equals 277.27384 Cu. In.	In Bushels equals 8 Gallons equals 2218.19075 Cu. In.
Millilitre, or cubic centimetre.....	0.061027	0.0000353	0.001761	0.00022010	0,000027512
Centilitre, or 10 cubic centimetres	0.610271	0.0003532	0.017608	0.00220097	0.000275121
Decilitre, or 100 cubic centimetres	6.102705	0.0035317	0.176077	0 02200967	0.002751208
Litre, or cubic decimetre	61.027052	0.0353166	1.760773	0.22009668	0.027512085
Decalitre, or centistere	610.270515	0.3531658	17.607734	2.20096677	0.275120846
Hectolitre, or decistere	6102.705152	3.5316581	176.077341	22.00966767	2.751208459
Kilolitre, or stere, or cubic metre	61027.051519	35.3165807	1760.773414	220.09667675	27.512084594
Myriolitre, or decastere	610270.515194	353.1658074	17607 734140	2200.96676750	275.120845937

1 Cubic Inch equals 16.3861759 Cubic Centimetres. 1 Cubic Foot equals 28.3153119 Cubic Decimetres.
1 Gallon equals 4.543457969 Litres.

MEASURES OF WEIGHT.

	In English Grains.	In Troy Ounces equals 480 Grains.	In Avoirdupois Lbs. equals 7,000 Grains.	In Cwts. equals 112 Lbs. equals 734,000 Grains.	Tons equals 20 Cwt. equals 15,620,000 Grains.
Milligram...........................	0.015432	0.000032	0.0000022	0.00000002	0.000000001
Centigram	0.154323	0.000322	0.0000220	0.00000020	0.000000010
Decigram...........................	1.543235	0.003215	0.0002205	0.00000197	0.000000098
Gram................................	15.432349	0.032151	0.0022046	0.00001968	0.000000984
Decagram..........................	154.323488	0.321507	0.0220462	0.00019684	0.000009842
Hectogram	1543.234880	3.215073	0.2204621	0.00196841	0.000098421
Kilogram...........................	15432.348800	32.150727	2.2046213	0.01968412	0.000984206
Myriogram.........................	154323.488000	321.507267	22.0462126	0.19684118	0.009842059

1 Grain equals 0.064798950 Gram. 1 Troy oz. equals 31.103496 Grams 1 Pound Avd. equals 0.45359265 Kilogr.

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THE ELEMENTS.

COMPOUNDS.